CHEMISTRY FOR ORGANISED SCHOOLS OF SCIENCE

PARRISH AND FORSYTH

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CHEMISTRY

FOR ORGANISED SCHOOLS OF SCIENCE

BY

S. PARRISH, B.Sc., A.R.C.S. (LOND.)

CENTRAL HIGHER GRADE SCHOOL, LEEDS

WITH INTRODUCTION

BY

D. FORSYTH, M.A., D.Sc.

PRINCIPAL OF THE CENTRAL HIGHER GRADE SCHOOL, LEEDS

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INTRODUCTION.

WHAT should be included *now* within the scope of general education is a subject of much thought and discussion. On this matter there is much diversity of opinion even among practical educationists. The time is past for any scheme of education to be considered general, that advocates either a purely literary or a purely scientific training. The literary may render a student intimate with the great thoughts of the great minds of the world, and give him facility and power of expression; the scientific will give him patience in observation, neatness in manipulation, thoroughness in method, and the power to investigate. Both kinds of training are needed in a liberal education. As the inhabitants of earth this world must always be to men an object of interest, and our relations to it and its relation to us form subjects of even a more absorbing nature. The Science that reveals to us any of earth's secrets, or hidden stores, that have but waited for the skilled hand and trained mind to discover, must therefore be accorded a place in any scheme of general education. Whether it be from the curiosity—so important a factor in the youthful acquisition of knowledge—excited by its processes, or the painstaking devotion of its disciples, or the keenly correct method of its investigations which nothing but the very truth will satisfy, or the marvellous results of its discoveries, Science has paramount claims for educational consideration. The ordinary language of

every day, and the common occurrences and conveniences of civilised life render imperative the intelligent understanding of the elements of science. Indeed some knowledge of Science—its principles, methods of enquiry and results—are as much now the necessity of the educated *man*, as reading and writing are of the *child* of to-day.

Chemistry, as showing *the composition of things*, is a subject of great interest to young people who have acquired a sound elementary education, and whose minds are prepared by intelligent study for practical work and practical observation. From its methods chemistry commends itself to them (1) by experiment, by providing something for hands to do and for eyes to see, and (2) by gradually leading through this experimental work to the attainment of the knowledge of a principle.

These two methods form the special merits of this text-book. It shows what boys *and girls* (from 13 to 15 years of age) can do and understand in experiment and in theory. The book is the patient result of years of practical teaching and observation in young people in experiments, in demonstration-lessons, and of class-lessons. It shows the fruit of much thought and consultation, and the strict exclusion of what would be impracticable, useless, or irrelevant. No experiments are included in it but have been practically done in the laboratory, and yet alternative experiments afford variety to the lessons, and add suggestiveness to the teaching of them. The abundance of exercises and questions is intended to present the meaning of the experiments in as many different aspects as possible.

The intimate relation between physics and chemistry forms a special feature in the book, and is well worthy of the notice of teachers. To students who have not had the advantage of practical training in a physical laboratory this should be of great advantage, while to those who have had some training

in physics the connection will be refreshing as well as interesting.

The utmost simplicity of language has been sought, consistent with clearness of explanation and strictness of definition. The illustrations for the experiments are in every case, where necessary, *fresh*, and made from actual apparatus.

The results that have been gained by such a two-years' course in this school may be fairly and briefly stated under two heads:

A. By intelligent procedure from the simple to the difficult, from the known to the unknown, from experiment to principle a basis is laid for future attainment in this subject by pupils showing special ability or inclination.

B. Sufficient is done in it to produce an appreciation of scientific method, a habit of scientific observation, and scientific accuracy and manipulative skill—and therefore to warrant the inclusion of such a course in the curriculum of the general education of those not seeking for more special attainments in chemistry.

D. FORSYTH.

PREFACE.

THE method of treatment adopted in this book follows, to a great extent, that employed in the two-years' Course in Chemistry in the School of Science in the Leeds Higher Grade School.

While it is not possible in a Course of Elementary Chemistry to claim any great advantage for one order of treatment over many others that might be suggested, my experience has satisfied me that the order adopted in the following pages, is both reasonable and practical. The early consideration of carbon and its compounds is, perhaps, unusual, but on the whole this is justified by the fact that carbon compounds afford so many interesting and instructive exercises, and require so many and varied processes of manipulation, all of which are well within the capabilities of students at the end of the first year.

In the first year, commencing with the phenomena of burning, students acquire a working knowledge of the air and the part it plays in processes of combustion. The balance is at an early stage introduced to their notice, but only *alterations* in weight are at first to be noted. Solutions are next considered, and this affords abundant opportunity for varied practice in manipulation. The solution of metals leads naturally to the study of hydrogen, after which oxygen, water, oxidation and reduction, and the formation of salts are considered. The first years' work ends with the study of carbon, its oxides, and a few organic compounds.

In the second year, a larger proportion of the experiments are quantitative. Chemical theory, symbols, formulae, etc., are not introduced till near the middle of the course, and even then with due consideration. The student's attention is throughout directed to the fact that chemical principles are founded on the result of previous experiment, his own results being used as far as possible, to make this clear.

Experiments introduced by "Expt. —" are in all cases laboratory experiments, and have been performed over and over again by young students of both sexes; they may all, therefore, be regarded as *possible* experiments. They are, almost without exception, such as can be completed in a lesson of 1½ hours duration, an important matter when many classes of students pass through the laboratory each week. Alternative experiments are frequently suggested, and these serve a double purpose: (1) they make it possible to give a greater variety to the teaching, and (2) they give an opportunity to specially bright pupils, to utilise their time to the full.

In order to afford some idea of the degree of accuracy that may be expected, I have given, in an appendix, some results, taken at random from students' note-books, of experiments described in the text.

The great majority of the experiments can be performed by the students working individually. This ensures a proper amount of practice for each student, while the chief advantage of working in pairs, viz., comparison and discussion of results need not be lost. Results arrived at independently by two students, are at least as useful for purposes of discussion as those which are obtained by a combined effort. When two pairs of hands are more effective than one pair—and this is notably the case in preparations—then working in pairs is both advisable and necessary.

Unexpected difficulties continually arise in a chemical labora-

tory in the performance of apparently straight-forward experiments ; teachers will, I hope, be able to derive some assistance from the hints which are frequently given in the following pages. At the same time, I should feel extremely grateful for any recommendations or suggestions that teachers might be inclined to offer.

Although this book was written without regard to any examination syllabus, the method of treatment adopted follows very closely that suggested in the new and revised syllabus for Elementary Chemistry in the Science and Art Directory just published (Aug. 1899). I trust, therefore, that the work will prove of service to teachers of evening classes.

I have to thank Mr. H. Smith for his kindness in photographing various pieces of apparatus for new diagrams, and numerous friends—teachers and others—for much assistance in the way of criticism and advice.

S. P.

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CHEMISTRY FOR ORGANISED SCHOOLS OF SCIENCE.

CHAPTER I.

THE METRIC SYSTEM.

The Unit of Length.

THE system of measurement employed in Chemistry is that known as the Metric System. The unit of length in this system is the **metre.** It was originally intended that a metre should be $\frac{1}{10{,}000{,}000}$ of a quadrant of the earth. The circumference of the earth would then be 40,000,000 metres. More accurate measurements show that the value of the metre is slightly different from this theoretical value, and now a metre is defined as the distance between two marks on a certain bar of platinum in the possession of the French Government.

1 metre = 39.37 in., very nearly.
= 1.093 yds. „ „

If we wish to convert measurements in yards into measurements in any other English unit, *e.g.* into feet, inches, or miles, we have to multiply or divide by various awkward numbers, 12, 3, 1760, etc. In the metric system the values of the fractions and multiples of a metre are:

1 decimetre $= \frac{1}{10}$ metre.	1 dekametre = 10 metres.
1 centimetre $= \frac{1}{100}$ „	1 hectometre = 100 „
1 millimetre $= \frac{1}{1000}$ „	1 kilometre = 1000 „

Reduction of lengths in the metric system therefore involve only multiplying or dividing by 10, or some multiple of 10.

The line in the figure is just 1 decimetre (or 3.937 in.) long, and is divided into 10 centimetres. Each centimetre is further divided into 10 millimetres.

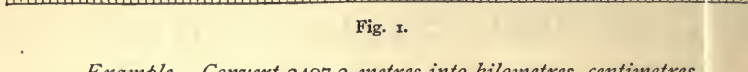

Fig. 1.

Example. *Convert* 2497.2 *metres into kilometres, centimetres, millimetres.*

1 kilometre = 1000 metres ;

$\therefore$ 2497.2 m. $= \frac{2497.2}{1000}$ km. = 2.4972 km.

1 metre = 100 centimetres ;

$\therefore$ 2497.2 m. = 2497.2 × 100 cm.

= 249720 cm.

1 metre = 1000 millimetres ;

$\therefore$ 2497.2 m. = 2497.2 × 1000 mm.

= 2497200 mm.

As 1 kilometre = 1000 metres, and 1 metre = 1.093 English yards, 1 kilometre = 1093 yards, nearly. Measurements which are in English expressed in miles, are expressed in the metric system in kilometres.

The Unit of Volume.

The unit of volume in the metric system is related to the unit of length in the following way.—Imagine a vessel made in the form of a cube, with each edge of the interior 1 decimetre long. This vessel would hold the unit volume, or 1 **litre.**

1 litre = 1.76 pints, nearly.

The multiples and fractions of the litre are named in the same way as the multiples and fractions of the metre.

Thus 1 dekalitre = 10 litres.
1 decilitre $= \frac{1}{10}$ litre.
1 millilitre $= \frac{1}{1000}$ litre.

We do not usually apply the term millilitre to this last quantity.

1 cubic decimetre = 1 litre.

1 decimetre = 10 centimetres ;

$\therefore$ 1 cubic decimetre = 10^3 cubic centimetres.

$\therefore$ 1 litre = 1000 cubic centimetres.

The term cubic centimetre, usually written c.c., is used instead of the name millilitre.

That 1000 c.cs. = 1 litre may be seen in the following way.—Draw a square 1 dm. edge. This may be divided into 100 squares of 1 cm. edge. To cover this square with cubes of 1 cm. edge, 100 would be required, and these would form a layer 1 cm. high. Ten such layers, consisting of 1000 small cubes, would be required to give 1 cubic decimetre. Hence we see that

1 litre = 1000 cubic centimetres.

The litre and the c.c. are the units of volume generally used in Chemistry.

The Unit of Weight.

The unit of weight is called the **gram**. A gram is the weight of 1 c.c. of distilled water at the temperature of its maximum density, *i.e.* at 4° C.

1 gram = 15.432 grains.

As 1 lb. avoirdupois = 7000 grains,

$\therefore$ 1 lb. = $\frac{7000}{15{\cdot}432}$ grm. = 454 grm.

The other weights employed are related to the gram in the following way :

$\frac{1}{10}$ gram	= 1 decigram.	10 grams	= 1 dekagram.
$\frac{1}{100}$ „	= 1 centigram.	100 „	= 1 hectogram.
$\frac{1}{1000}$ „	= 1 milligram.	1000 „	= 1 kilogram.

1 kilogram = 1000 grms., and 454 grms. = 1 lb.;

$\therefore$ 1 kilogram = $\frac{1000}{454}$ lbs. = $2\frac{1}{4}$ lbs., nearly.

This is the only unit greater than 1 grm. in general use.

QUESTIONS. I.

1. What is the length of 1 ft. in centimetres; of 1 mile in kilometres; of 1 decimetre in inches?

2. What are the dimensions in mm. of a test tube, 6 in. long, and $\frac{3}{4}$-in. diameter?

3. How many mm. in 1 in.?

4. A glass tube is of $\frac{5}{8}$-in. bore, what is this in mm.?

5. How many yards in 71 kilometres?

6. How many cubic millimetres in 1 cubic centimetre; in 1 cubic decimetre; in 1 cubic metre?

7. How many cubic mm. in 1 cubic in.?

8. How many gallons in 1 cubic metre?

9. If a pint of water weighs 20 oz., what will be the weight in lbs. of 1 cubic metre of water?

10. A cubic foot of water weighs 62.5 lbs.; how many litres are there in a cubic foot?

11. How many mgrm. in 1 grm.; in 1 kilogram?

12. How many grams, decigrams, centigrams, and kilograms in 10 lbs.?

13. Express in English weights and measures 500 grms., 500 c.cs., 500 mms.

14. Sulphuric acid is 1.8 times as heavy as water, while linseed oil is .835 times as heavy; find the weight of 75 c.cs. of each of these substances.

15. A solution of ammonia is 0.88 times as heavy as water, what volume of this solution will weigh 50 grms.?

16. Mercury is 13.6 times as heavy as water, what weight of mercury will fill a bottle which holds 135 grms. of water?

17. Find the weight in grms. of a piece of iron, 10 cm. long, 2 cm. broad, and 7 cm. thick, if iron is 7.8 times as heavy as water.

18. Define the gram and litre. How many grams are contained in 1 725 kilograms?

19. Find the weight of a block of sulphur 50 c.cs. in volume and of relative density 2.05.

CHAPTER II.

EFFECTS OF HEATING METALS.

Introductory.

The Bunsen Burner. In a great many of the experiments described in the following pages, the application of heat is required. A "bunsen" burner is very generally employed in the laboratory for heating purposes ; the student should therefore examine it carefully, and accustom himself to using it in the proper manner.

EXPT. 1. Notice that when the holes at the bottom of the tube are open, a smokeless flame is produced, which has not the same appearance throughout. An inner cone is very clearly marked. Introduce the head of a match into this inner cone, and hold the match steady. The wood takes fire before the head.

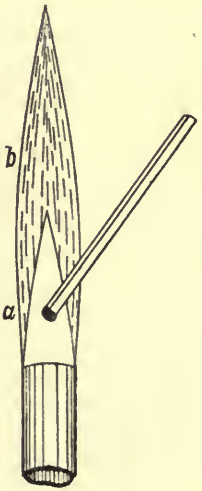

Fig. 2.

EXPT. 2. Bring a sheet of white paper, held parallel with the bench, smartly over the flame, at about the level of the point marked *a*. Hold it there for a few seconds—not long enough for the paper to take fire—and note the result. What does it prove ?

EXPT. 3. Bring another sheet into the flame in the same manner, at about the level of *b*. Compare the result with that obtained in Expt. 2.

It will now be obvious that if we wish to heat a substance, we should not hold it in the inner zone, but rather in the part of the flame marked *b*.

EXPT. 4. Hold a piece of narrow glass tubing in the bunsen flame. It becomes hot without being blackened with soot. This is the chief advantage of using a bunsen burner. Close the holes at the bottom of the tube of the bunsen. A yellow flame is produced, which soon blackens the glass tube.

EXPT. 5. Hold one end of the glass tube used in Expt. 4, in the inner cone, and apply a light to the other end. A small gas-flame is obtained. What does this prove?

EXPT. 6. Close one of the holes at the bottom of the bunsen with the finger, and bring near the other hole a narrow strip of smouldering brown paper. The smoke is deflected, and passes through the hole into the tube. This shows that, as the coal gas passes up the tube, it draws air through the holes. It is to be remembered, however, that the air drawn in in this way is not sufficient to burn all the gas.

To discover the effect of heat on various metals.

Before commencing any experiment, the materials to be used must be carefully examined, and a description written.

EXPT. 7. Heat a piece of platinum wire, held by a pair of "crucible tongs," in the bunsen flame. It becomes red-hot and gives out light as long as it is held in the flame. On taking it out it very soon becomes cold, and, *as far as we can judge by its appearance*, has returned to its original state.

EXPT. 8. Clean a piece of thin copper wire with emery cloth. Hold the wire in the flame as in Expt. 7. The copper glows as the platinum did while in the flame, and a globule of molten metal forms on the end. On taking the wire out, it is found that the part which was in the flame is covered with a dull black film, which, for the present, we may call "copper tarnish." Rub with emery again. We see that the tarnish is only on the surface.

EXPT. 9. Clean a piece of magnesium ribbon. On heating this, it burns, giving a dazzling white light. White fumes are

given off, and there is left a dull white ash, which retains the shape of the original metal, and is easily powdered.

The effect of heat on iron wire, and on tin, lead, zinc, all of which may be held in small iron spoons, should be noticed. In all these cases a tarnish is produced.

When the metal, as for example, lead, melts, if the tarnish is scraped off as it forms, the clean shining metal can be seen, and by continuing to expose fresh surfaces the whole of the metal may be changed.

We shall next inquire more closely into the changes which occur in experiments such as the above. And first we shall find out whether any change in *weight* occurs, whether the tarnish weighs as much as, more than, or less than the original metal.

We must now therefore become acquainted with the chemist's most important instrument, the balance.

CHAPTER III.

THE BALANCE.

The Balance is simply a pair of scales, made and adjusted so carefully as to show very small differences in weight of two substances.

The balance shown in Fig. 3 is found to be not too elaborate for beginners from twelve to fifteen years of age. If it is kept closed when not in use, it may be allowed to stand in the laboratory, on a table specially provided for the purpose. The injury it suffers is far less than it would incur by being continually carried away to a cupboard. In the earlier experiments, weighings correct to 1 cgrm. are sufficiently good ; afterwards, results correct to 1 or 2 mgrm. may be obtained.

The author finds that balances such as the above, after continual use by many classes of young students for four years, will still indicate a difference of 1 mgrm.

The beam is supported in the middle by a wedge of hard

steel, or of agate—a "knife-edge"—resting in a very shallow groove, also of steel. A similar arrangement is used for supporting the scale-pans, but in this case, the knife-edge is

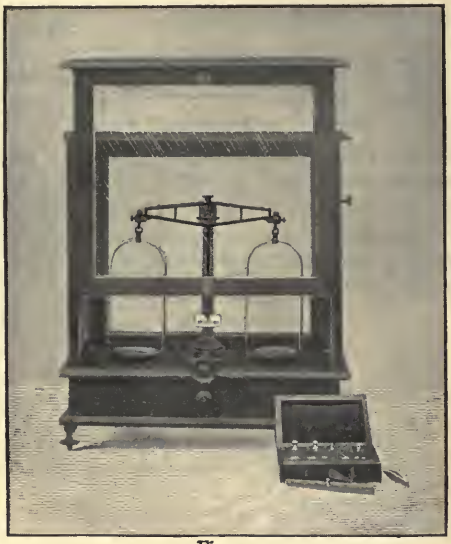

Fig. 3.

on the end of the beam. The steel should be protected by a very thin coating of vaseline.

By turning the screw placed outside the balance-case, the beam may be raised so as to allow it to swing, or lowered so as to prevent any motion. When not in use it should always be lowered.

A pointer is fixed to the middle of the beam, and when the beam is swinging, the end of this pointer moves over a white graduated scale. When the two pans balance, the pointer will move over the same number of divisions on each side of the zero position.

The weights to be used range from 50 grm. to 1 milligram.

The weights below 1 cgrm. may be made of aluminium wire. *Each weight should have a separate place in the box.*

The weights are arranged as follows:

50 grm.	5 grm.	0.5 grm.	0.05 grm.	0.005 grm.
20 „	2 „	0.2 „	0.02 „	0.002 „
10 „	2 „	0.1 „	0.01 „	0.002 „
10 „	1 „	0.1 „	0.01 „	0.001 „

Rules to be observed in weighing.

(*a*) Put the weights on the right-hand pan of the balance.

(*b*) Never put anything on the balance-pans, or take anything off, while the balance is free to swing.

(*c*) Always use the forceps provided, for lifting the weights.

(*d*) On commencing to weigh, find a weight which is too great, then, after removing this, try the succeeding weights *in order*. *Never pick out weights at random.*

(*e*) Do not put the small weights in a heap. Arrange them in order round the larger weights, which should be in the centre of the balance-pan.

(*f*) Place yourself opposite the centre of the graduated scale while weighing.

(*g*) Do not remove any weight from the balance-pan until the values of all have been written down, and check your result as the weights are replaced.

(*h*) Be careful to put the weights back in their proper places.

(*i*) Never attempt to weigh anything which is not quite cold. In addition to injuring the balance, the weighing will not be accurate.

The student should examine the balance, noticing the manner of working, and after a demonstration of the method of weighing, he should, for practice, weigh several objects, a watch-glass, coins, etc., and have the results verified before any of the weights are removed.

It is impossible to prevent the small weights being frequently lost. The five, two, and one milligram weights can be made from aluminium wire (B.W.G. No. 30) by the teacher, with very little difficulty. A piece about 30 cm. long is weighed, and the length which goes to 1 mgrm. is calculated. A piece of the wire is then

stretched on a board between two nails, and, using a vernier microscope, scratches are made with a needle-point to mark lengths corresponding to the weights required. The wire can then be cut with a fine sharp chisel.

CHAPTER IV.

EFFECTS OF HEATING.—*CONTINUED.*

EXPT. 10. **To find whether any change in weight occurs when platinum is heated.**

Weigh a piece of clean platinum wire on a watch-glass. Heat the wire for a few minutes, allow to cool, and weigh again. If the second weighing does not give exactly the same result as the first, the difference should be so very small as to justify our saying the weight is unaltered. Compare the result of Expt. 7, Chap. II.

EXPT. 11. **To find whether any change in weight occurs when copper is heated.**

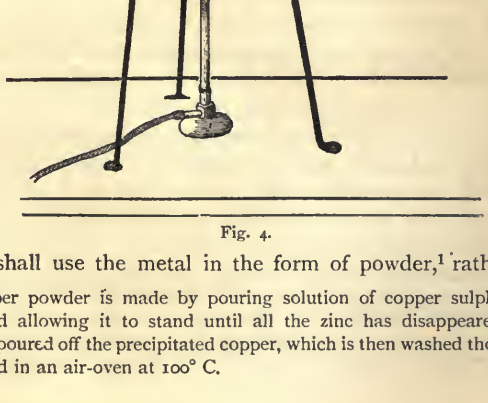

Fig. 4.

We shall use the metal in the form of powder,[1] rather than

[1] Copper powder is made by pouring solution of copper sulphate over zinc, and allowing it to stand until all the zinc has disappeared. The liquid is poured off the precipitated copper, which is then washed thoroughly, and dried in an air-oven at 100° C.

the wire. Put a little in a porcelain crucible provided with a lid, and weigh. Then heat the crucible over the bunsen in the manner shown in Fig. 4, raising the lid occasionally, for about 10 minutes. Allow to cool, and weigh again. A decided increase in weight will be observed.

EXPT. 12. **To find whether any change in weight occurs when magnesium is heated.**

The experiment will be carried on as in Expt. 11. About 6 ins. of clean magnesium ribbon should be used. An increase in weight is again noticed.

The effect of heating other metals, e.g. *lead, tin, zinc, iron, or aluminium, should be tried. The method employed in the two preceding experiments may be used.*

Platinum and a few other metals are exceptional in not suffering any permanent change on being heated; the great majority show a decided change, which corresponds in each case to an increase in weight.

To show that an increase in weight occurs when phosphorus is burnt, the following experiment may be made.[1]

A piece of hard glass tube (Fig. 5), about 6 in. long, is drawn out at one end. Near this, at *D*, a piece of phosphorus, about the size of a pea, is placed. Towards the wider end, from *B* to *C*, the tube is loosely packed with ignited asbestos fibre. The tube is then weighed, and connected with the arrangement shown in the figure, whereby air can be drawn through the tube. On gently warming the phosphorus, it takes fire and burns. The supply of air is to be so regulated that none of the white fumes made by burning the phosphorus escapes from either end of the tube *AB*. When all the phosphorus is burnt, the tube is

[1] It is not advisable to allow young beginners to attempt this experiment. It may be conveniently performed by the teacher in a demonstration lesson. In every case where an experiment is described in the following pages without "Expt. —" prefaced, the same direction is to be understood.

allowed to cool, and then weighed. A decided increase will be observed.

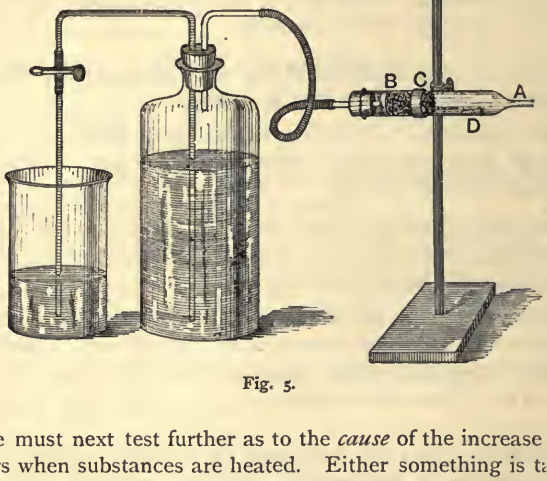

Fig. 5.

We must next test further as to the *cause* of the increase that occurs when substances are heated. Either something is taken from the *flame* or else from the *air*. The following experiment will throw some light on this question.

To show that when phosphorus burns in air, the air loses in weight just as much as the phosphorus gains.

A piece of dry phosphorus is put into a dry round-bottomed flask of about 400 c.cs. capacity, the mouth of which is then tightly closed with a rubber stopper, and the whole is weighed. The flask is held well above a bunsen flame, and rotated slowly, the object being to heat the flask as uniformly as possible, till the phosphorus takes fire. The flask is then held at arm's length, still being rotated, over a trough containing water, until the burning ceases.[1] At the end of the experiment

[1] If a small piece of asbestos cloth be put in the flask and the phosphorus be placed on it, the risk of the flask breaking is greatly diminished.

there should be some unburnt phosphorus left. It is then allowed to cool, and weighed again. *No increase in weight is observed*, consequently the mere act of heating does not cause any increase in weight. The mouth of the flask is then put under water, and the stopper removed. On shaking the flask about, the white smoke which is produced by the burning phosphorus soon dissolves in the water, which at once enters the flask. The fact that water enters shows that some of the air has been removed. The mouth of the flask is then covered with the hand, and the flask removed. If a lighted taper be introduced, it is at once extinguished. By measuring the total capacity of the flask, and the volume of water which enters, the fraction of the air which unites with the burning phosphorus is determined. This is found to be about one-fifth of the whole.

In burning, therefore, the phosphorus unites with something from the air, and this is the cause of the white solid weighing more than the original phosphorus. Only part of the air can unite with the phosphorus, the greater part, four-fifths of the whole, being an inactive gas which will not allow phosphorus or a taper to burn in it. As there is no *nett* gain or loss in the above experiment, the gain in weight of the phosphorus is just equal to the loss in weight of the air.

For measuring the capacity of any vessel, it will be found convenient to fill it with water, and to measure the volume of this by means of a measuring cylinder, such as that shown in the figure. Cylinders of 50 c.cs. and 200 c.cs. capacity will be found most useful. Notice in which direction the cylinder is graduated, and the volume between successive marks. In making a reading, place the cylinder on the bench, have the eye on a level with the surface of the water, and take the reading of the mark opposite the bottom of the curve made by the surface of the water.

Fig. 6.

The determination of the fraction of the air removed by phosphorus may also be conveniently made by allowing the phosphorus to smoulder, as in the following experiment.

EXPT. 13. **To determine with what fraction of the air phosphorus can combine.**

Twist the end of a piece of thin copper wire into a small loop. Put a small piece of phosphorus into a test-tube, cover it with water, and melt the phosphorus by putting the test-tube into a larger one containing hot water. Place the copper wire in the tube so that the loop is in the melted phosphorus, and cool it by pouring cold water on the outside of the tube. Push the end of the wire on which the phosphorus is fixed more than half-way up a tube closed at one end, of $\frac{1}{2}$ in. bore, and about 18 in. in length, and *at once* put the open end in a small beaker or wide-mouthed bottle containing water. The phosphorus slowly unites with part of the air, white smoke is produced, and during a lesson of ordinary length the water will have risen appreciably in the tube. Put the tube aside till next lesson. The water will rise to a certain point, and there remain stationary, although unchanged phosphorus may still be seen. Without taking the end of the tube out of the water, take out the copper wire, and at once put the phosphorus under water. Cover the open end of the tube, take it out of the water, turn it mouth upwards, and test with a lighted splint. This is at once extinguished. Now measure the volume of water in the tube, and so determine how much of the air has been removed. Next fill the tube with water, measure the volume of this, and so obtain the total volume of air used.

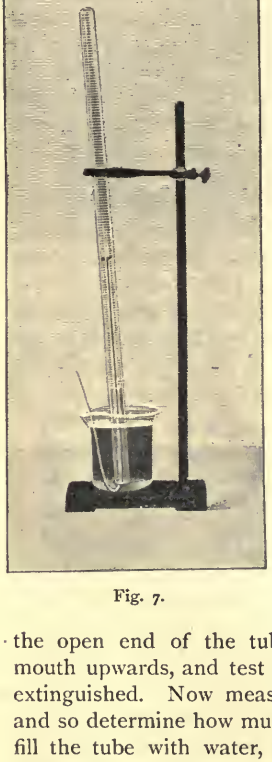

Fig. 7.

It is found, in every successful experiment of this kind, that

about one-fifth of the original volume of the air is removed by the phosphorus; the residual inert gas is about four-fifths of the whole.

We shall next compare this result with the action that goes on when metals tarnish. We shall commence with the metal iron. This tarnishes rapidly, particularly when damp, in air, even at the ordinary temperature.

To determine with what fraction of the air iron combines in rusting.

A little water and some fine iron filings are placed in a glass cylinder, and shaken until the inside of the cylinder is covered with the filings. It is then put, mouth downwards, into a vessel of water, and left in a warm place. In a few hours it will be seen that iron rust has begun to form, and the water has risen in the cylinder. The height of the water should be marked from time to time. In a day or two the water will cease to rise. The mouth of the cylinder should then be covered with a glass plate, the cylinder removed from the trough, and the gas tested with a lighted taper. As in previous cases, the light is extinguished. Further, it will be found on measuring the volume of water which has entered, and the capacity of the cylinder, that very nearly one-fifth of the air has disappeared. This suggests that the tarnishing or rusting of iron is a similar action to that which goes on when phosphorus burns.

When metals tarnish, either at the ordinary temperature, or when heated, or when other substances—for example, phosphorus—are burnt, the changes that go on are of precisely the same kind. In each instance, part of the air—about one-fifth of the whole—unites with the metal or other body, and consequently the product weighs more than the original substance. It will have been noticed that it differs also in other respects. Thus, from the white shining metal, magnesium, which can be drawn into wire, or rolled into thin ribbon, we obtained a white earthy-looking substance, without lustre, and which can easily be reduced to powder between the fingers.

Whenever a change is brought about which results in the

formation of a substance, or substances, different in composition as well as in properties from what we start with, we call it a **chemical change**. Most of the experiments of this chapter afford examples of chemical change.

Lavoisier's Experiments on Air.

The true explanation of what goes on in burning, and when metals are tarnished, or "calcined," as it was formerly called, was first given by a French chemist, Lavoisier.

In one of Lavoisier's experiments he heated mercury in a measured volume of air. An apparatus such as he used is

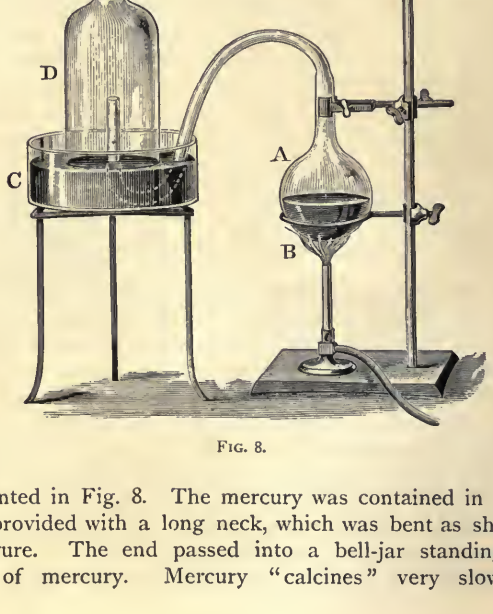

FIG. 8.

represented in Fig. 8. The mercury was contained in a glass globe provided with a long neck, which was bent as shown in the figure. The end passed into a bell-jar standing in a vessel of mercury. Mercury "calcines" very slowly on

heating.[1] Lavoisier heated it to near its boiling point for about 12 days. The volume of air got less and less, but at the end of that time no further diminution was taking place. The diminution was found to be nearly one-fifth. The red powder, or "calx," that had formed on the mercury, was collected and heated strongly (Fig. 9).

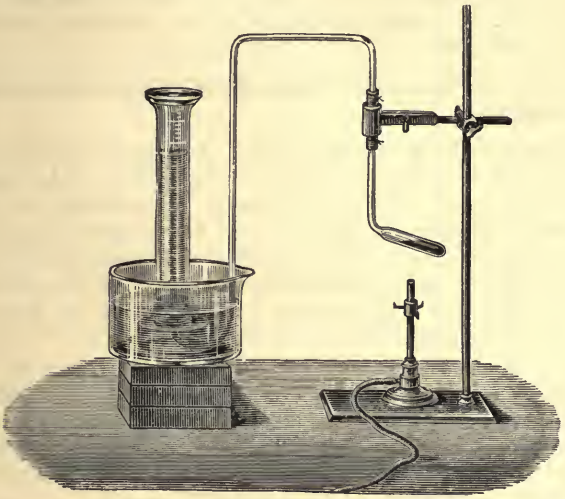

Fig. 9.

A gas was evolved which was just equal in volume to that which the air had lost, and which, when added to the residual gas, gave ordinary air again. Lavoisier also showed that the gas left when mercury was calcined in a limited supply of air would not allow substances to burn in it. We obtained the same result when iron rusted, and when phosphorus burned. By testing

[1] The formation of small amounts of mercuric oxide from metallic mercury may be shown in the following way: A small flask containing a little mercury is provided with a cork through which passes a thermometer reading to 300° C., and a glass tube, about 12 in. long, to act as an air-condenser. After heating for a few days on a sand-bath to 250°-300°, quite appreciable amounts of mercuric oxide may be seen on the surface of the mercury.

the gas obtained by heating the "calx" of mercury strongly, he found that it allowed substances to burn in it very energetically.

It was thus proved, therefore, that air is composed of two parts, one of which—about four-fifths of the whole—is very inert, the other—about one-fifth of the whole—is very active. It is with this latter part that substances unite in burning. By mixing the two constituents in the above proportions, ordinary air is again produced.

Expt. 14. To observe the effect of heating the calx of mercury.

Heat a little of the red "calx" of mercury, or mercuric oxide, in a test-tube, covering the mouth of the tube lightly with the finger. Notice the change in appearance of the powder. Introduce a splint of wood with a glowing end into the tube from time to time. Presently it bursts into flame, and burns brightly. A grey film of mercury is seen on the cold sides of the tube, and if the heating is continued, all the mercuric oxide will disappear.

The gas in the air with which burning substances unite is called **oxygen.** The products are called **oxides.** Substances which burn in air, burn much more energetically in oxygen. The inert gas, which does not allow bodies to burn in it, is called **nitrogen.** Its function in the air seems to be to dilute the oxygen.

QUESTIONS. IV.

20. A piece of charcoal is heated in the air until it disappears leaving only a white residue (ash). Give an explanation of this change.

21. Describe an experiment to show that a substance when burnt in air gains in weight. How do you account for this increase in weight?

22. You are supplied with pieces of the metals copper, platinum, magnesium, lead; how would you proceed to distinguish them from one another?

23. Describe all that you see on heating mercuric oxide.

24. A piece of copper wire, heated in air, becomes coated with a dull black substance. How do you account for this?

25. Phosphorus is a substance which is sometimes seen as a yellow waxy-looking solid, and sometimes as a red powder. Suggest a means of proving that these are one and the same substance.

26. How would you show that about one-fifth of the air is oxygen?

27. Give three examples of chemical change.

28. Why does a candle not continue to burn in a closed bottle?

29. Describe how Lavoisier proved that the atmosphere is composed of oxygen and nitrogen.

30. What difference in (*a*) appearance, (*b*) weight, is produced when the following are heated in air:

1. Mercuric oxide.
2. Platinum.
3. Copper.

CHAPTER V.

SOLUTIONS. DISTILLATION. PROPERTIES OF WATER.

EXPT. 15. Shake up a little common salt in water. The salt ceases to be visible. It is said to be **dissolved** in the water. The water is called a **solvent,** and the resultant liquid is a **solution** of salt in water. Has the salt ceased to exist? Taste the solution. The salt can still be detected. Put the solution in a porcelain dish and heat it. Soon the water boils away or evaporates, but the salt is left.

This is what usually happens when a solution of a solid in a liquid is heated.

EXPT. 16. **To find whether any change in weight occurs when a solid dissolves in water.**

Put about 10 c.cs. of water into a beaker small enough to be covered by a watch-glass. Put a little common salt or sugar on the watch-glass, and then weigh the whole. Carefully drop the solid into the water, and shake the beaker gently till the solid is dissolved. Replace the watch-glass and weigh again. No change in weight is observed.

EXPT. 17. **To prepare a saturated solution.**

Shake a little powdered nitre or saltpetre in water. It dissolves. Add more nitre, a little at a time, shaking well after each addition. Can the water dissolve an unlimited amount?

A solution in which there is as much solid dissolved as possible is called a **saturated solution**.

Now warm the solution with the undissolved nitre. What happens? Add more nitre and warm again. What do you learn from these experiments with regard to the solubility of nitre in water? Let cold water run over the outside of the tube containing the hot solution. Nitre separates in the form of fine powder. Does this verify your previous conclusion?

Now warm the tube again, taking care to dry the outside first. Once more a clear solution is obtained. Put the tube aside to allow the solution to cool slowly. Nitre again separates. Is the appearance the same as before? Instead of a fine powder, we obtain crystals of considerable size. Select a few of the most regular shape and draw them.

When a solid, in forming, takes a definite shape, bounded by regular sides or faces, it is said to crystallize. The body so formed is a **crystal**. What is seen in the case of nitre applies generally; the more slowly crystals form, the larger they are. Does all the nitre separate when the hot solution is cooled? Find out by trial in the next experiment.

EXPT. 18. Pour the liquid off the crystals into a clean evaporating dish, and heat as in Expt. 15. A further quantity of nitre is obtained. A little thought will convince the student that the liquid from which the crystals separated must have been *saturated* for its own particular temperature.

EXPT. 19. **To test whether Plaster of Paris is soluble in water or not.**

Fill a test-tube to the depth of about 2 in. with water, put into it a little plaster of Paris, say as much as will lie on a sixpenny-piece, and shake. The appearance does not tell us whether anything has been dissolved or not—some solid is still observed in the liquid. To find whether any solid has been dissolved or not, we must first separate the liquid from that solid

which is still undissolved (or is in *suspension* in the water). This might be done by allowing the solid to settle, and pouring off the clear liquid, but a much quicker method is to filter it.

A demonstration of the method of folding the filter-paper is here required. After the paper is dropped into the funnel, wet it, and press it into contact with the glass all the way round. Never try to fold a filter-paper with wet hands.

Having obtained a clear liquid by filtering, or a **filtrate**, as it is called, boil it to dryness in a porcelain dish. A residue is left, but if the volume of liquid used was about equal to that used in Expt. 18, it will be seen that the amount of residue obtained is much less than in the case of the nitre.

EXPT. 20. Repeat Expt. 19, but using powdered chalk instead of plaster of Paris. A still smaller amount of residue will be obtained.

The solubility of other substances may be tested, e.g. sand, charcoal, flour, marble, washing-soda, potassium chlorate, alum.

Solution in other liquids.

Although water will dissolve more or less of the majority of solids, yet many will dissolve to a greater extent in other liquids. A few examples of the use of other solvents may be given.

EXPT. 21. Shake a little powdered sulphur with (*a*) water, (*b*) alcohol, (*c*) carbon bi-sulphide. Note the result in each case.

EXPT. 22. Try similar experiments with shellac, rosin, saltpetre, iodine, or charcoal.

Arrange the results in the form of a table.

SUBSTANCE.	WATER.	ALCOHOL.	CARBON BI-SULPHIDE.
Sulphur.	Insoluble.	Insoluble.	Soluble.

If you are in doubt as to whether any appreciable amount of solid has dissolved, *e.g.* in the case of sulphur in water, filter a

few drops into a watch-glass, and evaporate on a sand-bath. To evaporate carbon bi-sulphide, put the watch-glass in a draught cupboard, and let the liquid evaporate spontaneously.

DISTILLATION.

We have seen that when water containing a solid in solution is boiled, the solid is left behind, while the water passes off as steam. If this steam is cooled, it changes again to water, and this water will obviously be pure.

EXPT. 23. **To prepare pure water from impure water.**

Half fill the bulb of an 8 oz. retort with a measured volume of ordinary water, rendering it evidently impure by

Fig. 10.

adding a crystal of copper sulphate, or blue vitriol. Pass the end of the neck of the retort into a 4 oz. flask which should res

on a shelf—the "bee-hive shelf"—in a "pneumatic trough." The trough should previously have been filled with water to the depth of two or three inches. Heat the water in the retort, and when steam begins to come off, cool it by pouring the water from the trough, by means of a basin, over the flask. The steam changes into water again or **condenses**. The first portion passing over will have washed out the neck of the retort, and after rinsing it round the inside of the flask, pour it into a measuring glass. Then continue the experiment until about half the water has boiled away, and measure the volume of liquid obtained. When it is cold, measure also the volume of liquid left in the retort. The sum of the three volumes will be found to agree very approximately with the original volume of water used.

An operation such as the above is called **distillation**, and the liquid that is condensed is called the **distillate**. The distillate, being colourless, is certainly free from copper sulphate.

The change from liquid water to the colourless gas, steam, and the further change from steam to water again, on cooling, afford simple examples of what is called a **physical change**. By this term we mean a change in appearance or other properties, without any change in the actual material, or *stuff*, we are dealing with.

In another method of distilling, the steam is cooled by passing through a tube surrounded by a wider one, the space between being kept full of cold water. This is a "Liebig condenser" (Fig. 11). The liquid to be distilled is conveniently placed in a "distilling flask," a round-bottomed flask with a narrow side-tube blown on the neck. Notice that the water for cooling a Liebig condenser enters at the lower opening. What would be the effect of allowing it to enter at the top?

When larger quantities of distilled water are required, the condenser takes the form of a spiral, or worm, *D* (Fig. 12), usually made of tin, coiled in a vessel of cold water. The flask is replaced by a copper boiler, *B*.

EXPT. 24. **To demonstrate the presence of solid in solution in ordinary drinking-water.**

Place a shallow glass dish, *e.g.* a clock-glass of about 4 in·

diameter, on a water-bath similar to that used in Expt. 32. Pour in water to within half an inch of the edge, and allow it to

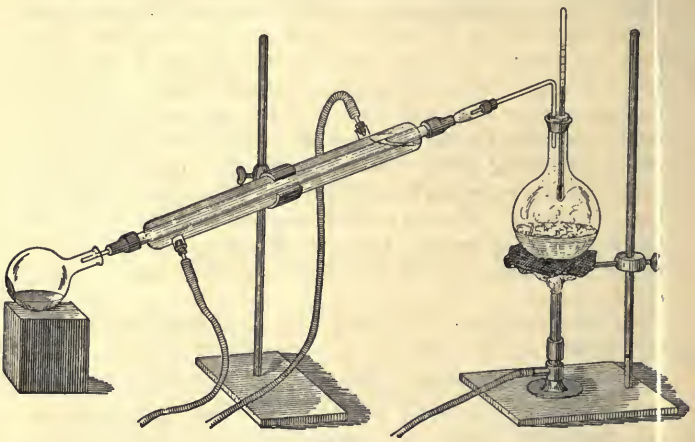

Fig. 11.

evaporate. The residue is rendered much more apparent in this way than if a porcelain dish were used.

EXPT. 25. **Distilled water leaves no residue on evaporation.**

Put a little distilled water in the glass dish, and evaporate to dryness. Practically no residue is left.

As a rule, the amount of solid dissolved in drinking-water is so small that beginners should not be asked to try and determine the amount An experiment may be made by the teacher, and the result entered in the students' note-books. From 70 c.cs. to 500 c.cs. of the water, according to the quality, should be used. A dish, 4 in. diameter, made of platinum, glass, or aluminium, is cleaned and weighed. The sample is then evaporated on the water-bath. The dish should not be filled further than within $\frac{1}{2}$ in. of the edge.

When all the water has disappeared, the dish is put in an air-oven at 105° C. for half an hour, and after being allowed to cool in a desiccator is weighed.

Distilled water is "flat" or "dead" to the taste, possessing

none of the refreshing qualities of spring water. In the next experiments we shall find out the cause of this difference.

EXPT. 26. Half fill a test tube with ordinary water and

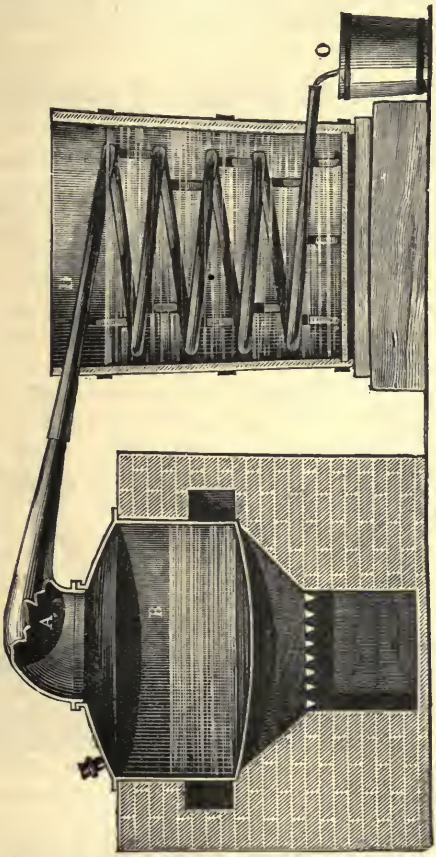

Fig. 12.

warm it. Bubbles of gas appear on the sides of the tube long before the water boils. Continue the heating. The bubbles

gradually rise to the surface and escape. Boil the water for a few minutes, then allow it to cool, and taste it. It will be found to have the taste of distilled water. It is evident therefore, that merely boiling for a short time (all the dissolved solid being retained) is sufficient to impart the "flat" taste to water.

To demonstrate the presence of dissolved gas in drinking-water.

A large flask (Fig. 13) is filled completely with water. The exit tube, one end of which must pass just through the stopper,

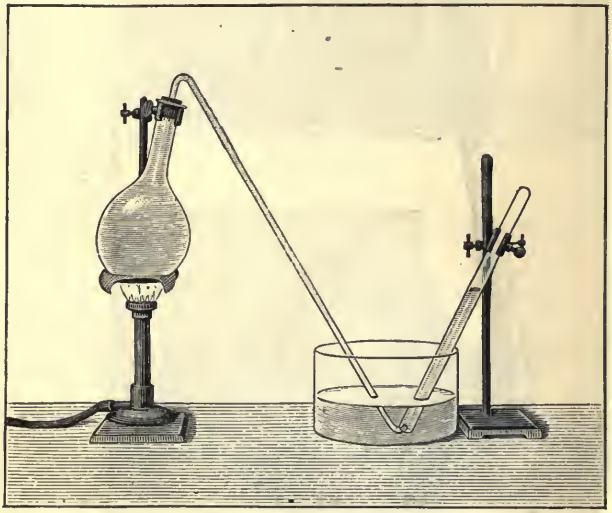

Fig. 13.

is also full. The other end passes under water in a pneumatic trough, and over this end is placed a graduated tube. When the water in the flask is boiled, the steam is condensed in

passing through the water in the trough. A certain amount of gas is not so condensed, however, and this collects in the measuring tube. Here its volume can be read off, and gives us, approximately, the volume of gas that was dissolved in the water originally used. All natural waters contain a certain amount of gas in solution, and it is this dissolved gas that causes spring water to taste so different from distilled water. By shaking boiled water with air for a little time, some of the air is dissolved, and the water acquires the taste of ordinary spring water.

Natural waters. Evaporation from the surface of the sea is continually going on. When the vapour gets cooled down, it is re-converted into water, and falls as rain. **Rain water** is, therefore, really distilled water. When collected in country districts it contains practically nothing in solution but a little gas dissolved from the atmosphere. In towns it always contains a certain amount of solid dissolved from the impurities floating in the air. **River water** and **spring water** always contain more or less solid in solution, the amount depending on the nature of the rock with which the water has been in contact. **Sea water** contains a great amount of solid, chiefly common salt, in solution. Waters to which the substances in solution impart a distinct taste, are called **mineral waters.**

The weight of dissolved solids in some samples of water is given below:

	Grains per gallon.
Loch Katrine, - - - - -	1.96
Thirlmere, - - - - - -	3.57
Leeds water supply, - - - -	6.5
Thames (at Twickenham), - - -	22.4
Deep well (Thorparch, Yorks), - -	38.5
Old Sulphur Well (Harrogate), - -	1083
Sea water, - - - - -	about 2450

Properties of water. Pure water is tasteless, odourless, and when seen in small quantities, colourless, although in large quantities it appears blue. When evaporated to dryness it leaves no residue. It dissolves more or less of a vast number

of substances, and hence all natural waters leave a solid residue on evaporation. Gases also dissolve in water, though, as we shall afterwards discover, some are far more soluble than others.

We may at this stage, also, notice certain other physical properties of water.

The boiling-point of water.

EXPT. 27. Fit up an apparatus like that shown in Fig. 14. The test-tube, 1 in. wide, contains distilled water to the depth of about 1 in., and is provided with a cork bored with two holes. Through one passes a Centigrade thermometer, the bulb of which is to be placed, at first, in the water. Through the second hole a short glass tube passes, to allow steam to escape. Using a small flame, gradually heat the water. The mercury in the thermometer rises to a certain point, and then remains stationary, when the water commences to boil. This point will be found to be 100°, or very near that. The temperature at which water boils varies with the pressure of the atmosphere, but the extent of this variation at any given place is only slight.

Fig. 14.

EXPT. 28. Next raise the thermometer until the bulb is about ½ in. above the surface of the water. If the water is again boiled gently, the thermometer will now be heated only by the steam produced. The mercury will again rise until it reaches 100°, and then remain stationary. The temperature of the steam is the same as that of the boiling water.

EXPT. 29. Dissolve some potassium nitrate in water, and

boil the solution. First put the thermometer into the solution. The temperature will be found to be above 100° C. Continue the boiling for some time, so as to boil off some of the water. All the solid remains, and so we are really obtaining a stronger solution. The temperature indicated by the thermometer will gradually rise. Now take out the thermometer, allow it to cool, then wash the bulb carefully and replace it, so that the bulb is above the surface of the liquid, as in Expt. 28. The temperature indicated will be the same as when we heated the thermometer in the steam from pure water.

We see, then, that by dissolving a solid in a liquid, the solution has a higher boiling-point than the pure liquid. The more solid we dissolve in the same quantity of liquid, the higher is the boiling-point of the solution. The vapour passing off, however, is at the same temperature as that from the pure liquid. To determine the boiling-point of a liquid, therefore, the only safe way is to put the thermometer in the vapour given off on boiling. The point marked 100° on a thermometer scale is fixed, when the instrument is made, by putting the thermometer in the steam from water boiling under known conditions.

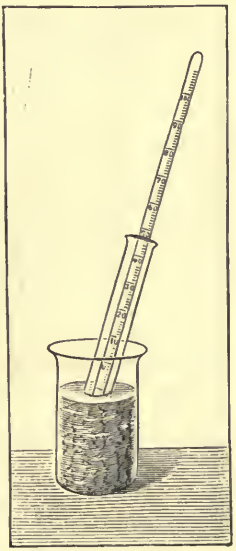

Fig. 15.

The freezing-point of water.

EXPT. 30. Put the bulb of a thermometer in a test-tube containing a little distilled water, and place the whole in a beaker containing a "freezing-mixture," made of snow, or ice shavings, and common salt, in the proportions roughly of 4 to 1. Stir the water gently with the bulb of the thermometer. Notice that the temperature falls until the water begins to freeze, when the thermometer will indicate 0° C. and then remain constant until all the

water in the test-tube is frozen. This is called the freezing-point of water.

EXPT. 31. When the water is frozen, take the tube out of the freezing-mixture and put it into a beaker containing water at the temperature of the laboratory. The ice begins to melt, but the thermometer does not rise above 0° C. until all the ice has melted. The constant temperature maintained during the time water is changing to ice, is the same as that maintained during the change from ice to water.

The point marked 0° C. on the thermometer is fixed when the instrument is made, by placing the lower part of the thermometer in melting ice, so that when the mercury comes to rest, the end of the column is just visible. The ice is usually placed, rather unnecessarily, in a vessel such as a funnel, from which the water as it forms can drain away.

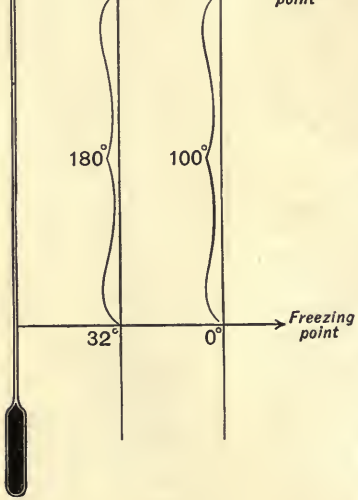

Fig. 16.

A description of the method of making a thermometer will be found in any text-book on Physics.

In a **Centigrade** thermometer, the space between the freezing-point and the boiling-point is divided into 100 divisions or degrees. This form of thermometer is most generally used in scientific work. In a **Fahrenheit** thermometer the freezing-point is marked 32° and the boiling-point 212°, so that the interval between these two fixed points is divided into 180°. The point marked 0° on a Fahrenheit thermometer is 32° below the freezing-point. 100

degrees Centigrade are equal to 180 degrees Fahrenheit, or

$$1^\circ \text{ C.} = \tfrac{9}{5}^\circ \text{ F.}, \qquad 1^\circ \text{ F.} = \tfrac{5}{9}^\circ \text{ C.}$$

To find what temperature F. corresponds to a certain temperature C. we first multiply by $\frac{9}{5}$. This gives us the number of degrees F. above the freezing-point. To find the number of degrees above the zero point, we must then add 32°. To change from the Fahrenheit scale to the Centigrade scale, we first subtract 32° and then multiply by $\frac{5}{9}$.

Example 1. *What temperature F. corresponds to* 50° *C.?*

$$1^\circ \text{ C.} = \tfrac{9}{5}^\circ \text{ F.}$$
$$\therefore \quad 50^\circ \text{ C.} = \tfrac{9}{5} \times 50 = 90^\circ \text{ F.}$$

That is, the required temperature is 90° F. above the freezing-point. To get the number of degrees above zero we must add 32°, therefore the required temperature is 122° F.

Example 2. *What temperature C. corresponds to* 180° *F.?*

180° F. is $(180 - 32) = 148^\circ$ F. above the freezing-point.

$$1^\circ \text{ F.} = \tfrac{5}{9}^\circ \text{ C.}$$
$$\therefore \quad 148^\circ \text{ F.} = (148 \times \tfrac{5}{9}) = 82\tfrac{2}{9}^\circ \text{ C.}$$

That is, the required temperature is $82\frac{2}{9}$° C. above the freezing-point. As the freezing-point on the Centigrade scale is 0°, the required answer is $82\frac{2}{9}$° C.

It is to be remembered that in changing from C. to F., the reduction is made first and 32 added to the result. In changing from F. to C., first subtract 32 and then reduce.

QUESTIONS. V.

31. Describe an apparatus by which you could prepare pure drinkable water from sea-water.

32. What occurs when nitre is placed in water? If the liquid is heated and more nitre added, what happens then? If now the liquid is poured off from any solid nitre that may remain, and is allowed to cool, what is the result?

33. State any facts known to you which serve to show that in the case of substances which are soluble in water more of a solid is dissolved by hot water than by cold water, and more of a gas is dissolved by cold water than by hot water.

34. In what respects do rain-water, river-water, and sea-water differ from each other?

35. How may the presence of air dissolved in water be demonstrated?

36. Black gunpowder is an intimate mixture of charcoal, sulphur, and nitre. Describe a method whereby these constituents could be separated from one another.

37. A colourless liquid is said to be water. Describe all the tests you know for proving whether this is so or not.

38. Given a mixture of sand and sulphur, how would you separate them from each other?

39. Mention three colourless liquids, besides water, you know. How do they differ from water? Which of those you mention have you proved to be good solvents for iodine?

40. What do you mean by a saturated solution? Describe exactly how to prepare a solution of potassium dichromate, saturated at the temperature of the laboratory?

41. If mud be shaken with water, what experiments would you make to ascertain whether the water has taken anything into solution?

42. In dissolving a solid in water, what is, in general, the result of (*a*) heating the solvent; (*b*) powdering the solid to be dissolved?

CHAPTER VI.

QUANTITATIVE EXPERIMENTS ON SOLUTION.

It will have been noticed that when a solution is evaporated nearly to dryness over a bunsen flame, it begins to spirt, and so we could not by this means obtain all the solid dissolved in a given liquid. The device shown in the next experiment is therefore generally employed.

Expt. 32. **To find the weight of solid dissolved in any given volume of a prepared solution.**

Weigh carefully a thin porcelain dish. By means of a narrow graduated cylinder measure out 20 c.cs. of the solution supplied. Half fill a beaker of the shape shown, and provided with a "lip"

or "spout," with water. The beaker must be of such size that the dish will rest easily on the rim. Put a few bits of paper in the water in the beaker; this will keep the water from bumping while boiling. Now place the dish in position, and as the water in the beaker boils, the steam will heat the dish, but never sufficiently to cause the water it contains to boil. By this means the solution is slowly evaporated without any loss of the solid. An arrangement such as the above is called a *water-bath*. When no more liquid is visible in the dish, wipe the outside, and place it for a few minutes on a *sand-bath*. This is merely a shallow tray containing fine sand, which is heated by a bunsen burner. By this means the solid is thoroughly dried, and *when the dish is cold* weigh it again. The increase in weight represents the amount of solid dissolved in 20 c.cs. of the solution.

Fig. 17.

It is possible, with care, to perform the whole evaporation on the sand-bath in experiments such as the above.

If not more than 20 c.cs. of solution are used, the above experiment can be completed in one lesson of 1½ hours' duration. Solutions of salts free from water of crystallization, *e.g.* lead nitrate, sodium chloride, potassium chlorate, potassium dichromate, should be used in experiments of this kind.

Expt. 33. To separate the constituents of a mixture of soluble and insoluble substances.

Make a mixture of nitre and powdered charcoal, shake with water, and then filter. The charcoal, being insoluble in water, is left on the filter-paper. Boil the filtrate to dryness. The potassium nitrate is again obtained as a white solid. The constituents of any mixture of a soluble and an insoluble substance can be separated in the same way.

For the next experiment mixtures containing known amounts

of a soluble and an insoluble constituent are required. Sand and salt, sand and sugar, chalk and sodium carbonate are suitable for this purpose.

EXPT. 34. **To determine the percentage of soluble and insoluble substances in a mixture.**

Fold a filter-paper, place it on the balance with a small dry porcelain dish, and weigh. Put about 2 grm. of the mixture in the dish, and weigh again. The increase represents the amount of mixture taken. Half fill the dish with water, and using a small glass rod, stir for a few minutes. Then let the insoluble part settle, and pour the liquid through the weighed filter-paper. Add more water, stir, allow to settle, and pour the liquid through the paper as before. To prevent the liquid running down the outside of the dish, pour it down the stirring-rod, which is held over the funnel. The succeeding quantities of water should be poured through the paper immediately after stirring. This will transfer the bulk of the insoluble part to the filter, and the remainder can be transferred by using a wash-bottle.

FIG. 18.

A demonstration of the manner of using a wash-bottle is required.

When all the insoluble substance is on the paper, wash it once or twice, and then dry it by means of the apparatus shown in Fig. 18. The "cone" is made of sheet-iron; air must have free access below, and the funnel may rest on a "pipe-clay triangle." The drying apparatus is to be heated on a sand-bath. When the paper is thoroughly dry, allow it to cool. Then put it in the dish, which must be clean and dry as at first, and weigh. The difference between the third weighing and the first gives the

weight of the insoluble matter in the mixture originally used. For purposes of comparison, all the results should be calculated to show the weight of each constituent in 100 grm. The percentage weight of the soluble matter is obtained by subtracting the percentage of insoluble matter from 100.

The results may be entered in the manner shown below:

Second weighing = 19.250 grm. = wt. of dish + paper + mixture.
First „ = 17.417 „ = wt. of dish + paper.
Difference = 1.833 grm. = wt. of mixture used.

Third weighing = 18.614 grm. = wt. of dish + paper + insol. part.
First „ = 17.417 „ = wt. of dish + paper.
Difference = 1.197 grm. = wt. of insoluble part.

	Grm. of Mixture.			Grm. of Insol. Part.
As	1.833	:	100	:: 1.197 : x

$$\frac{100 \times 1.197}{1.833} = 65.3\,\%\ \textbf{insoluble.}$$

$$\therefore\ 34.7\,\%\ \textbf{soluble.}$$

Definition of solubility. We have seen that different amounts of nitre dissolve in the same amount of water at different temperatures. The same remark applies to nearly all substances. The number of grams of a substance which will dissolve in 100 grm. of water at any temperature is called its **solubility** at that temperature. Clearly, a substance will have different solubilities, corresponding to different temperatures. The variation in solubility is well exhibited by a "graphic method."

Graphic representation of varying solubility.

Take two lines, AB and AC, at right angles. Suppose that we measure temperatures along AB, and weight in grams along AC, and that A corresponds to 0 grams and 0° C. If we agree that a certain distance along AB, say $\frac{1}{4}$ in., shall represent 10° C., then $\frac{1}{2}$ in. will represent 20° C., and so on. Next suppose the distances measured along AC represent grams of

solid dissolved in 100 grm. of water. Let $\frac{1}{4}$ in. measured along AC represent 10 grm.; then if we wish to represent 50 grm.,

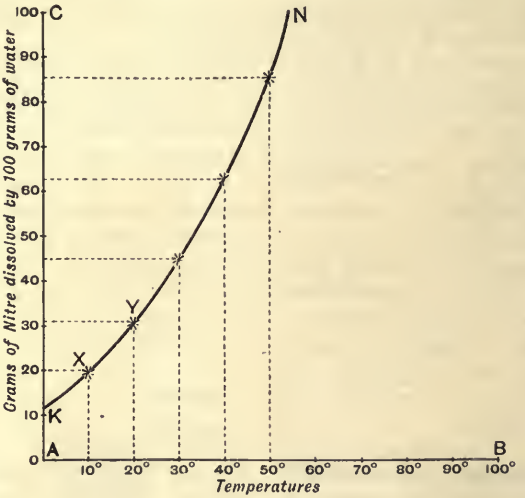

Fig. 19.

a point must be taken $\frac{5}{4}$ in. from A. Taking the case of saltpetre, it is found by experiment that the amounts dissolved by 100 grm. of water at various temperatures are as follow:

0° C.	-	12 grm.	30° C.	-	45 grm.
10° C.	-	20 ,,	40° C.	-	62 ,,
20° C.	-	32 ,,	50° C.	-	85 ,,

To show these facts on the diagram. Take, for example, the point on AB corresponding to 20° C. On AC take the point corresponding to 32 grm. From these points draw lines perpendicular to AB and AC respectively. Then Y, where these lines meet, indicates at the same time 20° and 32 grm. Corresponding points are found for solubilities determined at any other temperatures. By joining the various points found from experiment, a curve, called a **Solubility Curve**, is obtained,

from which the solubility at intermediate temperatures can be read off.

The solubility curves for several substances are shown in Fig. 20.

By using squared paper, such as is employed in a physical laboratory, the plotting of a solubility curve is very much simplified.

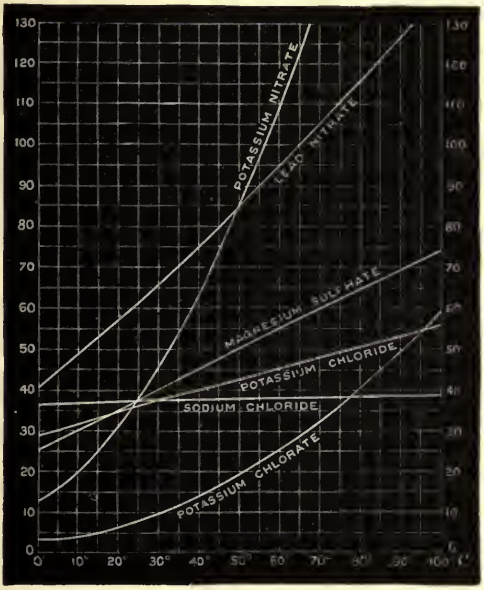

Fig. 20.

EXPT. 35. **To find the solubility of potassium chlorate at the temperature of the laboratory.**

Put a little powdered potassium chlorate into a large test-tube half full of water, and warm. When it is all dissolved, cool the solution under the tap. If any crystals separate, cover the

mouth of the tube, and put it aside till the following lesson. If no crystals are obtained, add more potassium chlorate, and repeat until some crystals separate, showing that the solution is then saturated. Then weigh a small porcelain dish, pour 15 to 20 c.cs. of the solution into the dish, and weigh again. Evaporate exactly as in Expt. 32, finish the drying on a sand-bath, allow to cool, and weigh. Enter results in the following manner:

Temperature = 17° C.

Third weighing = 20.110 grm. = wt. of dish + potassium chlorate.
First „ = 19.105 „ = wt. of dish.
Difference = 1.005 grm. = wt. of potassium chlorate.

Second weighing = 35.240 grm. = wt. of dish + solution (water + potass. chlorate.)
Third „ = 20.110 „ = wt. of dish + potassium chlorate.
Difference = 15.130 grm. = wt. of water.

	Grm. of Water.		Grm. of Potass. Chlorate.		
As	15.130	:	100	::	1.005 : x

$$\frac{1.005 \times 100}{15.130} = 6.6 \text{ grm.}$$

= wt. of potassium chlorate dissolved in 100 grm. of water.

The solubility of other substances, e.g. *potassium nitrate*, *sodium chloride*, *potassium dichromate*, *lead chloride*, *might be determined in the same way.*

QUESTIONS. VI.

43. In Fig. 20 it is seen that the solubility curve for sodium chloride is almost a horizontal line. What does this indicate?

44. The solubility curves for lead nitrate and potassium nitrate cross at the point corresponding to 51° C.; what is the meaning of this?

45. Find, from Fig. 20, the increase in solubility of potassium nitrate between 20° and 30°, between 30° and 40°, and between 40° and 50°.

46. How would you find the percentage of dissolved solids in sea-water?

47. River-water contains solids in solution and also in suspension.

How would you proceed to find the amount of each of these in 70,000 grams of water?

48. What is meant by the *solubility* of a substance? How would you proceed to determine the solubility of potassium dichromate?

49. Measuring temperatures along the vertical axis, and weights along the horizontal axis, make a solubility curve for barium chloride from the following data:

Temp.	Grams dissolved by 100 grms. water.	Temp.	Grams dissolved by 100 grms. water.
0	32.5	40	43
10	34.5	50	46
18	37	60	48
26	39	70	51
34	42	80	53.5

CHAPTER VII.

ACTION OF ACIDS ON METALS. HYDROGEN.

WE shall next study the manner in which metals may be dissolved. Common knowledge tells us water will not serve as a solvent in this case. We shall therefore try the effect of the following acids, viz. hydrochloric, sulphuric, and nitric, on a few of the common metals.

EXPT. 36. Pour dilute hydrochloric acid over a few pieces of granulated zinc[1] in a test-tube.[a] A brisk effervescence is immediately set up. Hold the finger loosely over the mouth of the tube for a short time, and then note the effect produced (*a*) on the gas, (*b*) on the splint, when a lighted splint is put in

[1] Granulated zinc is prepared by melting the metal in an iron ladle, and then pouring it, in a thin stream, from a height of a few feet, into cold water.

[a] This and similar notes refer to the equations at the ends of chapters.

the mouth of the tube. Notice also whether the gas causing the effervescence has any colour or odour.

In the same manner try the effect of hydrochloric acid on iron filings, magnesium, aluminium, tin, copper, and lead.

It sometimes happens that no apparent result is observed until the acid is warmed, and then the effect produced by warming the acid alone, is attributed to the action of the acid on the metal. This confusion may be avoided by first warming a little of the acid by itself, and then dropping in the metal.

If no effect is produced by the dilute acid, pour this away, and use a little of the strong acid.

EXPT. 37. Pour dilute sulphuric acid (1 volume of acid to 10 of water) over granulated zinc,[b] as in Expt. 36, and note the result. Also repeat with the various metals mentioned above.

In diluting sulphuric acid, always pour the acid, slowly, into the water, never the water into the acid, and stir well during the operation.

The action of strong sulphuric acid on metals had better be postponed till later (see p. 230).

EXPT. 38. Allow strong nitric acid, diluted with its own volume of water, to act on the same metals.

Now set out your results in tabular form.

METAL.	HYDROCHLORIC ACID.	SULPHURIC ACID.	NITRIC ACID.
Tin.	Dilute: no action. Strong, hot: colourless, odourless gas evolved, which burned, but extinguished the lighted splint. Colourless liquid left, which leaves a white solid on evaporation.	No result.	Brisk effervescence, brown gas with disagreeable smell evolved, which would not burn, and which extinguished the lighted splint. White powder left.

It will have been noticed that an inflammable gas is given off

in the majority of cases, when sulphuric or hydrochloric acid is used, but never when nitric acid is employed. This gas, formerly known as "inflammable air," we now call "hydrogen," for a reason that will be seen later (Chap. X.).

Preparation of Hydrogen.

We shall now prepare and collect a quantity of this gas, using zinc and sulphuric acid. For this purpose an apparatus similar to that shown in Fig. 21 must be arranged.

The delivery tube, conveniently made of two pieces joined by rubber tubing, will be easily made after a demonstration of the method of bending glass.

Points to be remembered in bending glass. Always remember

(*a*) to heat the glass tube *held lengthways* in an ordinary luminous flame.

(*b*) to keep it rotating gently while it is being heated.

(*c*) to prevent it from bending when you feel it is just getting soft. When you think it is sufficiently soft, take it out of the flame and bend it to the required shape.

(*d*) to smooth the ends of the tube, by holding in a bunsen flame until the glass softens.

Try to make your apparatus look as neat as possible. Avoid making the various branches of your delivery tube unnecessarily long, and see, by laying the tube on the bench, that the branches all lie in one plane.

Rubber stoppers, which can be bought ready bored, are to be preferred.

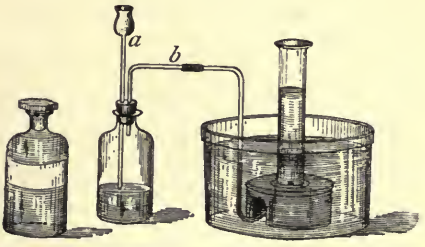

Fig. 21.

EXPT. 39. Put about 30 grms. of zinc in the generating bottle, allowing it to slide down the side. Cover the zinc with

water, and introduce the stopper, through which pass the thistle-funnel and the delivery-tube. See that the end of the funnel passes below the surface of the water. The end of the delivery-tube should pass under the bee-hive shelf, standing in the pneumatic trough. Put sufficient water into the trough to cover the shelf. Then pour a little strong sulphuric acid down the thistle-funnel. Bring an inverted test-tube filled with water over the hole in the top of the shelf. As the gas passes into the tube the water is expelled.

This method of collecting a gas is generally called the method of "displacement of water." It can only be used when the gas is insoluble, or only slightly soluble in water.

When the tube is full of gas close the mouth with the thumb, and bring it, *mouth downwards*, to a flame. Probably there is no result. Repeat this several times. Presently it will be found that there is a slight explosion, and later, that the gas takes fire and burns quietly.

Test-tubes, say 6 in. × $\frac{3}{4}$ in., are quite satisfactory for collecting the gas for the following experiments in the laboratory. The bee-hive shelf may then be dispensed with.

Properties of Hydrogen.

EXPT. 40. Place a tubeful of the gas, mouth downwards, in the stand, in such a way that the mouth is not completely closed. Leave it thus till the end of the lesson. On then testing with a lighted splint, it will be found that the gas in the tube will not burn at all.

The hydrogen has made its way downwards, and has escaped, while air has taken its place.

EXPT. 41. Pass a lighted splint into a tube of the gas held mouth downwards. The gas burns only at the mouth. What is the reason of this? Notice also that the splint is extinguished.

EXPT. 42. Try if you can pour invisible hydrogen from one tube to another. First try to pour it as you would pour water. Do you succeed? Then try in the following way:

Bring the tube, or small jar, containing hydrogen, marked *B* (Fig. 22), parallel to the empty jar *A*, which is held mouth downwards, and gradually turn *B* mouth upwards in the manner indicated. Apply a light to *A*, and notice the result. A slight explosion shows that hydrogen has been poured upwards from *B* to *A*, thus showing it to be lighter than air.

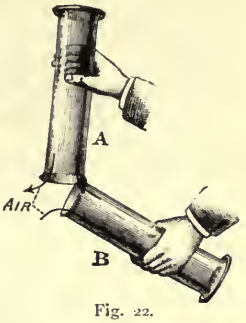

Fig. 22.

EXPT. 43. Arrange your apparatus with a delivery-tube shaped like that in Fig. 23. Hold an inverted test-tube over the end of the delivery-tube. In a short time it will be found that hydrogen free from air can be collected. This is known as the method of "upward displacement," and may be employed in collecting any gas lighter than air. It is particularly convenient when the gas is soluble in water.

Fig. 23.

EXPT. 44. [1] Blow soap-bubbles filled with hydrogen at the end

[1] To prepare the soap solution, dissolve a few grms. of sodium oleate in forty times its weight of cold distilled water, or of clean rain-water, and then add good glycerine equal in volume to about one-third or one-fourth that of the water. Allow to stand for a few days, then siphon the liquid off from any sediment or scum that may have formed, and add one drop of ammonium hydrate. Keep the bottle stoppered in the dark. Use a small quantity of the solution in a dish for experiments, and do not return any that may be left over into the stock bottle.

of an ordinary clay tobacco-pipe, or of a thistle-funnel. They rise rapidly through the air.

Similarly a small balloon may be inflated with hydrogen, and will rise rapidly to the ceiling.

That hydrogen is lighter than air may also be seen from the following experiment:

The beaker *B* (Fig. 24) is supported, mouth downwards, on

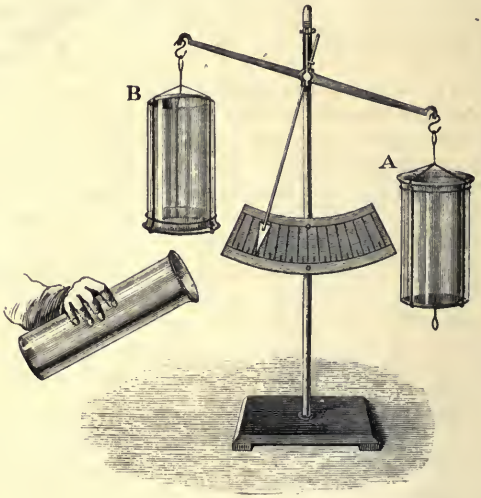

Fig. 24.

one arm of the balance, and counterpoised. On pouring hydrogen into the beaker, the movement of the index shows that the side on which is the beaker *B* has become lighter. Leave the beaker containing the hydrogen suspended for $\frac{1}{4}$ hour. At the end of that time, it will be found that equilibrium is again restored, showing that all the hydrogen has made its way out.

We learn, then, that hydrogen is a colourless, odourless gas, which burns in air, but does not allow a lighted splint to burn in it. It is almost insoluble in water. It is lighter than air; it is, in fact, the lightest substance known. In consequence of it

lightness we find, as in several of the experiments just described, that hydrogen has a great inclination to rise, and to rise rapidly, through the air. Nevertheless we see from the last experiment, and from Expt. 40, that it can also make its way downwards, although comparatively slowly. This point will again be referred to (see p. 128).

To liquefy hydrogen, a temperature of $-240°$ C. is required. This is lower than that required to liquefy any other known gas.

Hydrogen is taken as the standard of density for gases, with a relative density 1. When we say the density of oxygen is 16, we mean its relative density, that is, the number of times any volume of oxygen is heavier than the same volume of hydrogen.

Hydrogen is an example of a substance which has, so far, resisted all attempts to split it up into two or more simpler substances. Such substances we call **elements**. All the metals are elements. There are about 70 elements known at the present time; all other substances are made by the union of two or more of these elements. For a list of elements see p. 179.

VOLUME OF HYDROGEN OBTAINED BY DISSOLVING METALS.

EXPT. 45. **To find what volume of hydrogen is obtained when equal weights of different metals are dissolved.**

The apparatus shown in Fig. 25 will be found very convenient for making these experiments.

The tube *a* is easily made from soft glass tubing of about $\frac{5}{8}$ in. bore. From the wide end to the bend may be about 5 in. A piece of copper gauze is pushed in the wide end till it is about $1\frac{1}{2}$ in. from the mouth. Then the known weight of metal is introduced, and is kept in position by a second piece of gauze. The first piece of gauze prevents bits of metal from being carried up the tube during brisk effervescence, and there left

in a position which the acid does not afterwards reach. The delivery-tube *b* passes under the bee-hive shelf, and the gas is collected in the bottle *c*. A "kali" bottle, 20 oz. capacity, answers very well for this purpose.

When everything is ready, bring the beaker containing the solvent under the tube *a*. The tall, narrow form of beaker is best, and the wide end of the tube *a* should reach nearly to

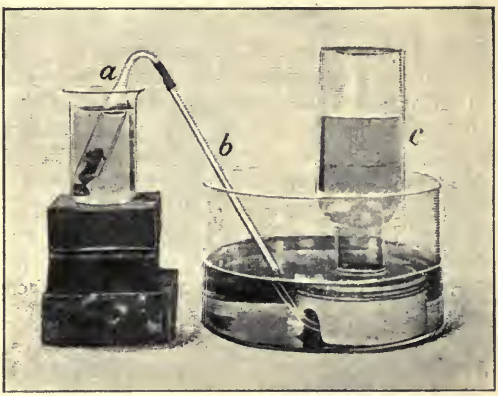

Fig. 25.

the bottom. If it is seen that the liquid does not rise in the tube *a* to reach the metal, pour more acid into the beaker, or take some water out of the trough.

When the metal is dissolved, mark the level of the water in the bottle *c* by a strip of gummed paper. Then fill the bottle with water to this mark, and measure the volume, using the measuring cylinder of 200 c.cs. capacity.

The following metals, with the solvents indicated, give satisfactory results.

Zinc: sulphuric acid diluted with seven times its volume of water.

Magnesium: (a) *sulphuric acid diluted with thirty times its volume of water;* (b) *hydrochloric acid diluted with seven times its volume of water.*

Aluminium: hydrochloric acid diluted with three times its volume of water. The acid should be warmed to start the action.

A solution of sodium hydrate (8 oz. sodium hydrate to 1 W. Qt. of water) may also be used to dissolve aluminium. This solution also requires warming.

The amount of metal used should be such as will yield 300 c.cs. of gas or thereabouts.

In order to compare results of various experiments, the volume of gas obtained from 1 grm. of metal should be calculated. It will be found that while 1 grm. of a certain metal always gives the same volume of gas, this is different from the volume obtained from 1 grm. of another metal. Moreover, it does not matter what the solvent is; 1 grm. of magnesium gives the same volume of hydrogen, whether it is dissolved in sulphuric or in hydrochloric acid.

The volume of gas obtained in these experiments is subject to various corrections which cannot now be entered into.

EQUATIONS.

(*a*) $Zn + 2HCl = ZnCl_2 + H_2$.

(*b*) $Zn + H_2SO_4 = ZnSO_4 + H_2$.

(*c*) $H_2 + O = H_2O$.

The student should ignore the equations until after Chapter XXVI. has been studied.

QUESTIONS. VII.

50. Describe the ordinary method of making hydrogen gas. Give a sketch of the apparatus used.

51. Give a summary of the chief properties of hydrogen.

52. Give experiments to prove that hydrogen is a very light substance.

53. Of two cylinders containing hydrogen, one is placed mouth upwards, and the other mouth downwards, and the cover slips are removed. What is the result in each case?

54. What is the origin of the name "hydrogen"?

55. What precaution must be observed in bringing a light to the end of the delivery tube of a hydrogen apparatus, and why?

56. Describe a method of collecting hydrogen without the use of the pneumatic trough.

57. What is meant by the term "element" in chemistry? What evidence is there that iron is an element?

CHAPTER VIII.

EXAMINATION OF RESIDUE LEFT IN THE HYDROGEN APPARATUS.

Water of Crystallisation.

It will be seen, on examining the hydrogen apparatus, that much of the zinc has disappeared; it has been dissolved. See if it is possible to recover it, as it was found possible to recover the salt from the solution in water, *i.e.* by evaporation.

Expt. 46. **To find whether zinc can again be obtained by evaporating the liquid left in the hydrogen apparatus.**

Filter the liquid left in the generating bottle; the filtrate is clear and colourless. Place it in a porcelain dish, and evaporate it over a small flame until a slight scum begins to form on the surface. Then put the dish aside to cool. A quantity of long, thin crystals, consisting of "zinc sulphate," separate, but no zinc is obtained. Pour off the liquid, and dry the crystals thoroughly between sheets of blotting-paper. In mass the crystals appear white, but the individual crystals are colourless.

Expt. 47. **To compare the weight of zinc sulphate with the weight of the zinc from which it is made.**

Weigh a small porcelain dish, then place in it about .2 grm. of zinc foil, and weigh again. Add dilute sulphuric acid until the zinc is dissolved, evaporate carefully to dryness on the water-bath, allow to cool, and weigh. A decided increase in weight will be observed.

To obtain zinc again from zinc sulphate, the crystals are dissolved in water so as to make a fairly strong solution. Two pieces of platinum foil, to each of which a piece of platinum

wire is attached, are placed in the solution about ½ in. apart. On connecting the wires to the terminals of a battery of four or five Grove's cells, or their equivalent, the current passes through the zinc sulphate solution, decomposing the zinc sulphate, and the metal zinc is deposited on the plate by which the current leaves the solution. On putting this plate into sulphuric acid, hydrogen is given off, and zinc sulphate is again produced.

In the experiment for the preparation of hydrogen, it was found that no action took place so long as the zinc was in contact only with the water. While the gas was being evolved, fresh quantities of acid had from time to time to be added. The acid became "spent," so to speak, as the zinc disappeared, and the liquid contained neither zinc nor sulphuric acid, but "zinc sulphate," as we have just discovered.

Zinc is a substance from which nobody has ever succeeded in obtaining anything simpler. The hydrogen must therefore have come from the sulphuric acid. Both the acid and the zinc ceased to exist as such, and we obtained instead the substance we have called zinc sulphate.

Water of crystallisation. EXPT. 48. Heat a few of the crystals of **zinc sulphate** in a dry tube, holding the tube with the mouth turned slightly downwards. The "dry" crystals give off a quantity of water, and the crystalline shape is lost, a white amorphous mass being left. The water which, as in this case, forms part of a crystal, is called water of crystallization.

EXPT. 49. Heat the remainder of the crystals in a porcelain dish till the water of crystallisation is given off. Allow to cool, then add a little water to the white mass. A good deal of heat is developed. Then heat with water sufficient to dissolve the whole, and allow to cool. Crystals are again obtained, exactly similar to those we obtained in Expt. 46.

EXPT. 50. Dry a few small crystals of **copper sulphate** with blotting-paper, then heat them in a dry tube. Water is again evolved, the crystalline shape is lost, and we also notice that the blue colour disappears. Add a few drops of water to the residue when cold. Heat is again generated, and the blue colour is restored.

EXPT. 51. Try whether **washing-soda, common salt,** and **potassium chlorate,** give off water on being heated.

Not all crystallised substances contain water of crystallisation. Where it does occur, we find that the crystalline shape is destroyed when the water is expelled. In the case of copper sulphate, we find that the colour also is lost. When water is added to the residue after heating, the heat produced is a sign that union with water is again taking place.

EXPT. 52. **To determine the weight of the water of crystallisation in 100 grms. of zinc sulphate crystals.**

Weigh a clean porcelain crucible and lid. Introduce about 1 grm. of the crystals, and weigh again. Now place the crucible with the lid on a cold sand-bath arranged as in Fig. 26. Place a lighted bunsen underneath, and heat for about 20 mins., raising the temperature very gradually. Remove the crucible, allow to cool, and weigh. Then heat again for 5 mins., cool, and weigh again. If the last weighing differs by a small amount only (less than 1 cgrm.) from the previous one, it may be regarded as correct

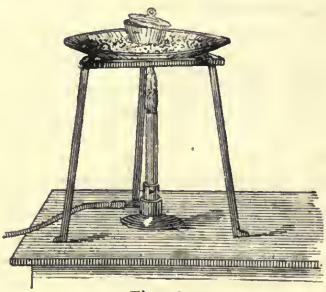

Fig. 26.

Enter results in the following way:

Second weighing = 20.411 grm. = wt. of crucible + crystals.
First " = 19.507 " = wt. of crucible.
Difference = 0.904 grm. = wt. of crystals.

Second weighing = 20.411 grm. = wt. of crucible + crystals.
Last " = 20.010 " = wt. of dish + (crystals − water).
Difference = 0.401 grm. = wt of water.

	Grms. of Crystals.			Grms. of Water.
As	0.904	: 100	::	0.401 : x

where x = percentage of water = 44.3 grms. of water.

The determination of the amount of water of crystallisation, i

any, in 100 grm. of copper sulphate, common salt, or magnesium sulphate might be made. The method of the last experiment is to be employed.

QUESTIONS. VIII.

58. Starting with metallic zinc describe exactly how to obtain a specimen of crystallised zinc sulphate.

59. A piece of zinc dissolves when put into hydrochloric acid, and a crystal of nitre dissolves when put into water. In what do the actions in the two cases differ from one another? How would you show that the actions were different in the two cases?

60. What do you mean by "water of crystallisation"?

61. Do all crystals contain water of crystallisation? If not, give examples of crystals which you know to be without it.

CHAPTER IX.

OXYGEN.

We found in Chapter IV. that when metals tarnish they unite with oxygen from the air, and make bodies called "oxides." We also found that when mercuric oxide was heated strongly, oxygen and mercury were obtained.[a] It might be thought that all oxides of metals would behave in the same way.

Expt. 53. Try the effect of heating copper oxide, yellow lead oxide, red lead, iron rust, manganese dioxide, or magnesium oxide, in dry test-tubes, as in Expt. 14.

Note the results, and in particular test for oxygen by means of a glowing splint. It will be noticed that red lead[b] and manganese dioxide[c] give off oxygen, while the others do not. When we require a quantity of oxygen, we do not obtain it from an oxide, but from a white solid called **potassium chlorate.**

Expt. 54. Heat a little potassium chlorate in a dry tube.[d] It first melts, and soon begins to give off bubbles of gas. Test

for oxygen as before. When the tube is cold, add a little water to the residue, and allow it to stand while the next experiment is made.

EXPT. 55. **To distinguish a chloride from a chlorate.** Dissolve in water, in separate test-tubes, a little potassium chlorate and potassium chloride. Add to each a few drops of a solution in water of the crystalline solid "silver nitrate." A white solid will be thrown down in the tube in which was the chloride, whereas in the case of the chlorate there will be no change. Add a little of the silver nitrate solution to solutions of other chlorides. In each case the result is the same. The formation of this white solid, or "precipitate," by means of silver nitrate, is a "test" for a chloride.

EXPT. 56. Add a few drops of silver nitrate solution to the liquid prepared in Expt. 54. A white precipitate shows that a chloride has been formed from the chlorate while oxygen was being evolved.

The preparation of oxygen.

EXPT. 57. Make a mixture of potassium chlorate with about one-third of its weight of manganese dioxide, and heat a little of it in a dry tube. Even before the potassium chlorate melts, oxygen is given off, as is shown by its rekindling a glowing splint. This mixture is the most convenient source of oxygen for laboratory use.

EXPT. 58. Bend a glass tube to the shape shown in Fig. 27. Put into a test-tube (6 in. × $\frac{3}{4}$ in.) sufficient of the mixture to half fill it. Have the mouth of the tube turned *slightly* downwards. *Take care that the mixture does not touch the stopper.* The oxygen will be collected by displacement of water, exactly as hydrogen was. Have ready five or six gas-jars of about 20 oz. capacity, filled with water, and covered with glass plates. Heat the mixture very cautiously by moving a small bunsen flame to and fro along the tube. Soon the gas begins to come off, and if the jars are placed in rapid succession on the shelf, from the quantity of material suggested there will

easily be obtained enough gas to fill them all. At first there will be a slight haze in the bottles, but on standing a short time this disappears, and the oxygen is seen to be colourless. It is without taste or smell. It is only very slightly soluble in water. Its

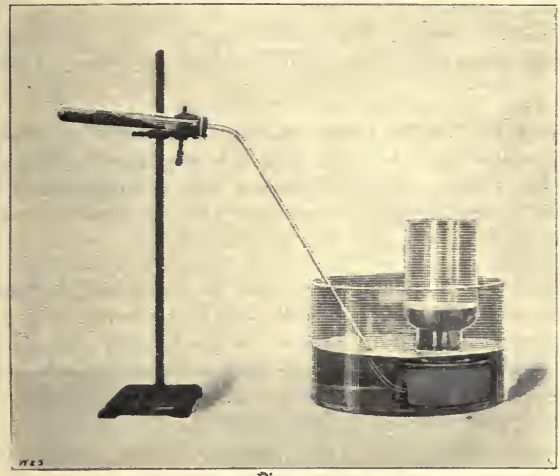

Fig. 27.

relative density is 16, *i.e.* it is, bulk for bulk, 16 times as heavy as hydrogen.

As we have found a wooden splint to burn brightly in oxygen, we may now try the effect of putting other burning substances into the gas. For the following experiments the deflagrating-spoon should be so arranged, that the little cup on which is put the substance to be burnt, is about 1 in. from the bottom of the bottle, and as far as possible from the sides.

Expt. 59. To burn a candle in oxygen.

Warm the cup of a deflagrating-spoon, and press a piece of candle into it. When cold, the candle will be securely fixed. Light the candle and put it into one of the jars. It is seen to burn much more brightly.

EXPT. 60. Put a small piece of **charcoal**[e] on a deflagrating-spoon. (A piece of charcoal formed from bark, or from a knot in the wood, is best.) Heat it till it begins to glow, and put it into one of the jars of oxygen. The charcoal burns much more brightly in the oxygen, and may send off a shower of sparks. Notice the appearance of the residue, or "ash," left in the spoon. Put a little water in the jar, cover the mouth with a glass plate, shake, and put on one side for future use.

EXPT. 61. In the same way, put a little burning **sulphur**[f] into another jar. It burns with a bright, blue flame, and the gas left appears somewhat hazy. Very cautiously notice the odour of the gas. Add water,[g] and place the bottle on one side as before.

EXPT. 62. Carefully dry a small piece of **phosphorus**[h] by means of filter-paper, put it in the cup of the deflagrating-spoon, which must also be dry, ignite the phosphorus, and introduce it into the next jar. It burns with a dazzling white light, a white cloud being formed at the same time. Add water[i] and shake. The white fumes soon dissolve.

Dry phosphorus very easily takes fire in the air. It should therefore never be touched with the dry hands.

EXPT. 63. Put a small piece of **sodium**,[j] which should be as clean as possible, into a deflagrating-spoon; heat it till it melts and takes fire, and introduce it into another jar. It burns more brightly. The residue, which will be white, or nearly so, and which will for the most part remain on the spoon, readily dissolves in water.[k]

Sodium must, on no account, be touched with wet hands (See Expt. 72.)

EXPT. 64. Introduce a piece of burning **magnesium**,[l] which may be conveniently held by the crucible tongs, into a jar of oxygen. It continues to burn, giving a dazzling white light, but the white solid produced is not, as far as we can judge by appearance, dissolved on shaking with water.

EXPT. 65. Put a little blue litmus solution[1] into each of the first four jars. It is turned red, the change in colour being less decided in the jars in which the carbon and the candle were burnt than in the other two. This property of turning blue litmus red is common to bodies called **acids**.

EXPT. 66. Pour a little of this reddened litmus into the jars in which the sodium and magnesium were burnt. Notice that in each case the blue colour is restored.

Those bodies which resemble the products of burning sodium and magnesium in their action on litmus, and in other respects, are called **alkalies**.

If another jar of oxygen is available, burn a piece of charcoal in it, and shake up the product with lime-water. This will become milky. The formation of this milkiness in lime-water is the best "test" for carbon dioxide, as the product obtained by burning charcoal is called.

To show the burning of iron in oxygen.

It is possible to get substances which do not burn in air to burn in oxygen. If a number of pieces of iron wire be rolled together, and a piece of filter-paper be rolled up, fastened to the end of the wire, and lighted, then, when put into oxygen,[2] the iron takes fire and burns readily, producing a shower of sparks. The oxygen for this experiment should be put in a common wide-mouthed bottle, the bottom of which is covered with wet sand.

In every case when a substance unites with oxygen, the product is called an **oxide**.

Some oxides, when dissolved in water, give acid liquids, which turn blue litmus red. Such are the oxides of carbon, called **carbon dioxide**, of sulphur, called **sulphur dioxide**,

[1] Litmus is a substance obtained from certain lichens. The solution in water is made red by acids, and blue by alkalies. Many other vegetable colours are changed by acids and alkalies. The infusion of red cabbage is a familiar example.

[2] Oxygen sold in steel cylinders is not always pure enough to allow this experiment to be performed successfully.

and of phosphorus, called **phosphorus pentoxide**. Other oxides, like **sodium oxide** and **magnesium oxide**, dissolve in water and give alkaline solutions, which turn red litmus blue. Still another class of oxides, of which the oxide of iron may be regarded as a type, do not dissolve in water, and so cannot be tested with litmus. The chemical relations of the various oxides will be considered in Chap. XIII.

Lavoisier thought that oxygen was a constituent of all acids—was, indeed, essential to the formation of acids. Hence the name "oxygen," given by him, and signifying "acid producer." Some acids, such as hydrochloric acid, are now known to contain no oxygen, and hence the name oxygen is a misnomer.

QUANTITATIVE EXPERIMENTS WITH OXYGEN.

EXPT. 67. **To find the weight of oxygen given off on heating 100 grms. of potassium chlorate.**

Weigh a clean dry test-tube, 6 in. long, made of hard glass. Introduce about 1 grm. of potassium chlorate, and weigh again. To ensure that none of the potassium chlorate remains on the sides of the tube, and so escapes heating, pour it through a roll of paper, reaching nearly to the bottom of the tube. Then heat the potassium chlorate, cautiously at first, and afterwards as strongly as possible in a bunsen burner. Allow to cool, and weigh.

Enter results in the following manner:

Second weighing = 22.416 grm. = wt. of tube + potassium chlorate.
First " = 21.890 " = wt. of tube.
Difference = 0.526 grm. = wt. of potassium chlorate.

Second weighing = 22.416 grm. = wt. of tube + potassium chlorate.
Third " = 22.214 " = wt. of tube + potassium chloride.
Difference = 0.202 grm. = wt. of oxygen evolved.

	Grms. of Potass. Chlorate.		Grms. of Oxygen.
Then as	0.526	: 100 : :	.202 : x

whence $x = 38.4$ grms. of oxygen.

EXPT. 68. **To find the weight of oxygen given off on heating 100 grms. of mercuric oxide.**

Weigh, as in Expt. 67, from 1 to 2 grms. of mercuric oxide, and heat cautiously. Take care not to heat the sides of the test-tube, else some of the liberated mercury may be driven out of the tube. When all the oxide has been decomposed, allow the tube to cool, and weigh again. Calculate as before.

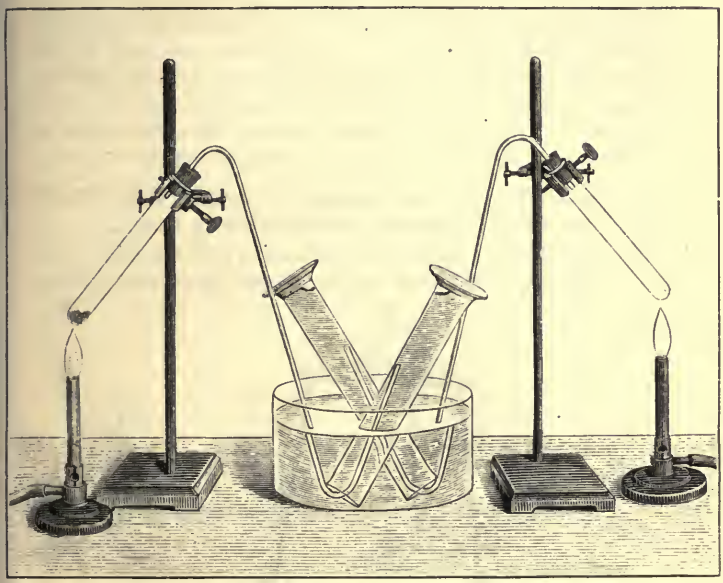

Fig. 28.

From 100 grms. of mercuric oxide we get very much less oxygen than from the same weight of potassium chlorate. As mercuric oxide is also much more expensive, it will be understood why potassium chlorate is preferred in the preparation of oxygen.

It is scarcely advisable for young students to try and determine the volume of oxygen evolved when potassium chlorate

and mercuric oxide are heated. The volumes obtained in the two cases may be conveniently compared in the following way (Fig. 28).

0.245 grm. potassium chlorate, and 1.296 grm. mercuric oxide, are carefully weighed into separate test-tubes of hard glass. These are fitted with delivery tubes, the ends of which pass nearly to the top of measuring cylinders of 100 c.cs. capacity, filled with water, and standing in a pneumatic trough. After heating, the tubes are allowed to cool down before the volumes of gas are measured. The test-tubes will then be full of gas at the same temperature as before heating, and the volume of gas in the cylinders will represent all that has been given off. It will be found that the very different weights of the substances used, give off equal volumes of oxygen.

EXPT. 69. **To find the weight of oxygen with which 100 grms. of magnesium combine when burnt.**

Weigh a small crucible, put in about .3 grm. of clean magnesium ribbon, and weigh again. Then heat the crucible

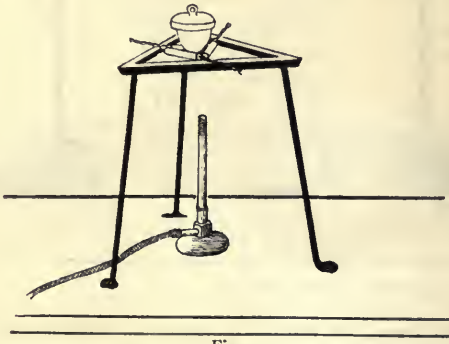

Fig. 29.

strongly for ten minutes with the lid on, then for another ten minutes, during which the lid is to be lifted slightly from time to time. After being heated for twenty minutes, the crucible should be allowed to cool, then weighed, heated for another five minutes, and weighed again. If the second weighing agrees

with the first it may be assumed that all the metal has been burnt.

Enter the result thus :

Second weighing = 15.427 grm. = wt. of crucible + magnesium.
First " = 15.153 " = wt. of crucible.
Difference = 0.274 grm. = wt. of magnesium.

Third weighing = 15.607 grm. = wt. of crucible + magnes. oxide.
Second " = 15.427 " = wt. of crucible + magnesium.
Difference = 0.180 grm. = wt. of oxygen with which the magnesium has combined.

Grms. of Magnesium. Grms. of Oxygen.
Then as 0.274 : 100 :: 0.180 : x

$$\therefore x = \frac{18.0}{0.274} = 65.7 \text{ grms. of oxygen.}$$

= weight of oxygen combined with 100 grms. magnesium.

The most accurate experiments show that 100 grms. of magnesium always combine with 66.6 grms. of oxygen.

EQUATIONS.

(*a*) $HgO = Hg + O$.
(*b*) $Pb_3O_4 = 3PbO + O$.
(*c*) $3MnO_2 = Mn_3O_4 + O_2$.
(*d*) $KClO_3 = KCl + 3O$.
(*e*) $C + O_2 = CO_2$.
(*f*) $S + O_2 = SO_2$.
(*g*) $SO_2 + H_2O = H_2SO_3$.
(*h*) $P_2 + O_5 = P_2O_5$.
(*i*) $P_2O_5 + 3H_2O = 2H_3PO_4$.
(*j*) $Na_2 + O = Na_2O$.
(*k*) $Na_2O + OH_2 = 2NaOH$.
(*l*) $Mg + O = MgO$.
(*m*) $Fe_3 + 4O = Fe_3O_4$.

QUESTIONS. IX.

62. Give a summary of the principal properties of oxygen.

63. Give two examples of metallic oxides which give off oxygen on heating, and two examples of oxides which do not.

64. Describe, with a sketch, the method you employed for filling a jar with oxygen.

65. By what means do you distinguish the products of burning carbon, sulphur, phosphorus, iron, in oxygen ?

66. Of two jars, one is known to contain oxygen and the other carbon dioxide. How would you distinguish between them ?

67. Describe all that happens when a crystal of potassium chlorate is heated in a dry tube.

68. Give two examples of acid-forming oxides, and two examples of alkali-forming oxides.

69. A white powder consists of either potassium chlorate or potassium chloride. Describe two experiments by which you could distinguish between these two substances.

70. Why does a candle burn more brightly in oxygen than in air?

CHAPTER X.

WATER.

WE shall now proceed to study the effects of burning hydrogen. We commence by burning a jet of hydrogen in a dry jar of air

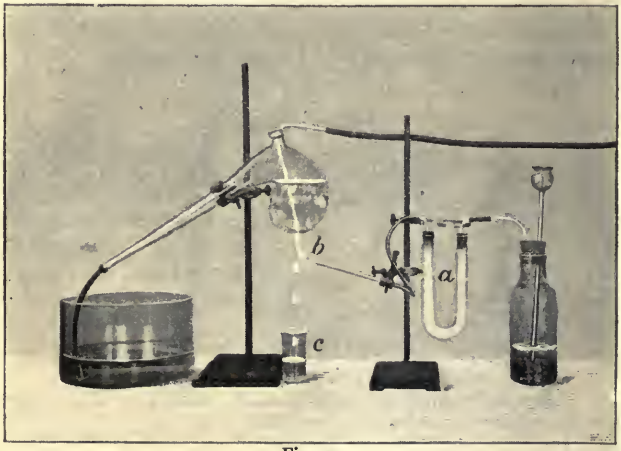

Fig. 30.

as shown in Fig. 31. We find that the sides of the jar become dimmed as though by moisture. This might be due to the burning of the hydrogen, or it might be moisture carried along with the hydrogen from the generating bottle.

In the next experiment, we shall find out whether, when hydrogen burns, water is really produced or not.

EXPT. 70. **To show that burnt hydrogen is water.**

Hydrogen, generated in the ordinary way, is passed through a tube *a*, containing pieces of fused calcium chloride (Fig. 30). This substance unites with water, and so will take out any moisture that may have been carried along with the hydrogen.

The dried gas is then burnt at a jet *b*, placed underneath a retort which is kept full of cold water. Drops of liquid collect on the outside of the retort, and fall into the beaker *c*, placed to receive them. When sufficient liquid has been obtained, it will be found to answer to all the tests for water. It will be without taste or smell, will leave no residue when evaporated, it will be neutral to litmus, its freezing-point will be 0° C., and its boiling-point will be 100° C.

The word "hydrogen" signifies "water-producer," and it will now be understood why the name was given.

To show hydrogen burning in oxygen.

A tube of the shape shown in Fig. 31 is connected to a tube delivering hydrogen which has been dried by passing through a calcium chloride tube. After the hydrogen has been escaping for a few minutes, it is ignited. On putting the tube into a dry jar containing oxygen, the hydrogen continues to burn, although the flame at first becomes considerably less. Why is this? The inside of the jar is seen to be dimmed as though by moisture.

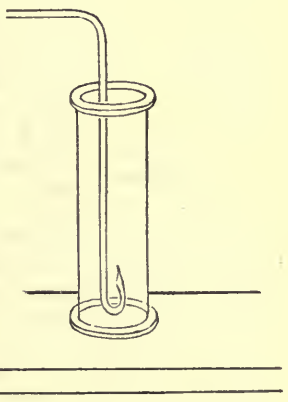

Fig. 31.

When hydrogen is to be burnt in this way, the end of the delivery tube in the generating bottle should be cut off at an

angle of 45°. *Then the drops of liquid which collect on it do not interfere with the continuous passage of the gas.*

Let us try next what would be the effect of applying a light to a jet of oxygen surrounded by hydrogen. The experiment may be made in the following way.

To show oxygen burning in hydrogen.

A dry jar is filled with hydrogen by upward displacement, and a light applied to the mouth. At once a tube delivering oxygen is pushed through the flame. The oxygen is seen to take fire, and continues to burn as long as there is any hydrogen with which it can combine. The appearance of moisture is again noticed. It appears, then, that it is a matter of indifference how the hydrogen and oxygen are presented to one another; if they are in contact and a light is applied, they will unite, and the flame is the visible sign that such union is going on. If a considerable quantity of hydrogen is in contact with oxygen, an explosion is produced when a light is applied.

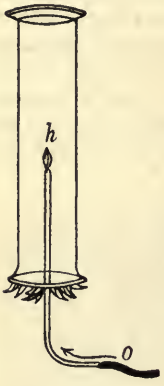

Fig. 32.

To explode a mixture of oxygen and hydrogen.

A soda-water bottle is filled, by displacement of water, with a mixture consisting of two volumes of hydrogen to one of oxygen. (It will afterwards be proved that these are just the proportions in which oxygen and hydrogen combine).

The bottle is then, for safety, wrapped in a duster, and, on applying a light, a flash is seen and a loud report heard. All the hydrogen unites with oxygen at once, a good deal of heat is produced, causing a sudden expansion of gas, and it is this that causes the report.

The oxyhydrogen flame. A flame of hydrogen fed with oxygen constitutes the oxyhydrogen flame. It is exceedingly hot, but gives out very little light. It is often used when a very high temperature is required, *e.g.* for melting platinum.

If oxygen instead of air is used in an ordinary "blow-through" blow-pipe, the "oxycoal-gas flame" which is obtained easily melts a piece of platinum wire.

The lime-light. When an oxyhydrogen, or oxycoal-gas flame is allowed to play on quicklime, the latter is not melted,

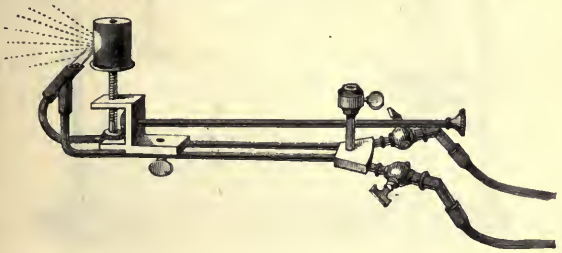

Fig. 33.

but becomes white hot, and emits a dazzling white light. This is the "lime-light" often used in magic lanterns.

Hydrogen from water.

Having proved that when hydrogen is burnt in air, water is produced, we shall next try to obtain hydrogen from water. This requires that we shall just reverse what we have done in previous experiments, *i.e.* that we shall separate the oxygen from the water. We have found that metals, as a rule, unite readily with oxygen, particularly when heated. We shall try, therefore, the effect of passing steam over heated magnesium or copper.

EXPT. 71. Pass steam from the flask *a* (Fig. 34) through the tube *cd*, in the bulb of which are heated, in turn, magnesium ribbon and copper turnings. On applying a light to the end of the tube, we see that we have a gas produced that will burn in air in the case of the magnesium,[a] but not in the case of the copper. The magnesium will be seen to be oxidized, while the copper is practically unaffected.

A broken test-tube, with the broken end almost sealed up, answers very well instead of the bulb-tube in this experiment.

A cheap and effective method of preparing hydrogen from water is to pass steam over red-hot iron. An iron tube *ab*,

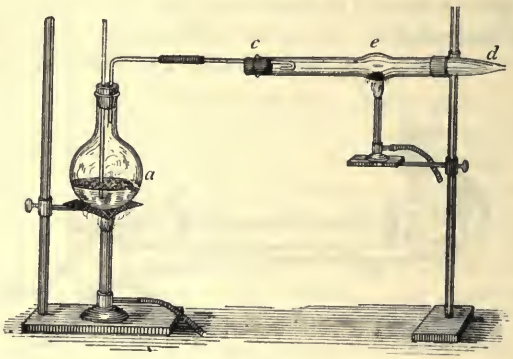

Fig. 34.

(Fig. 35), containing iron nails, is heated to redness in a furnace; steam is blown in at one end, *b*, and hydrogen is readily collected by displacement of water from the other. On examining the

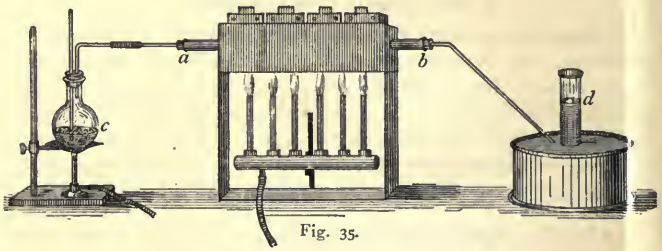

Fig. 35.

nails, they will be found to be coated with a grey substance, similar to that obtained by heating iron in air, as in Chapter II., or by burning iron in oxygen.

Some metals can liberate hydrogen from water, even at the ordinary temperature.

EXPT. 72. Throw a piece of sodium, the size of a pea, on to water contained in a porcelain dish. The metal swims about as

a bright globule, with a hissing noise, becoming less and less till it ultimately disappears. On rubbing the fingers together under the water, a soapy feeling will be noticed. Test the water with red litmus. It is turned blue. If successive small quantities of sodium be used in this experiment, and the solution evaporated, a white solid, called "**sodium hydrate**," is obtained.

To collect the gas evolved when sodium acts on water, the following method may be employed:—

EXPT. 73. Wrap a piece of sodium, the size of a pea, in wire gauze. Fill a test-tube with water, invert it in the trough in the

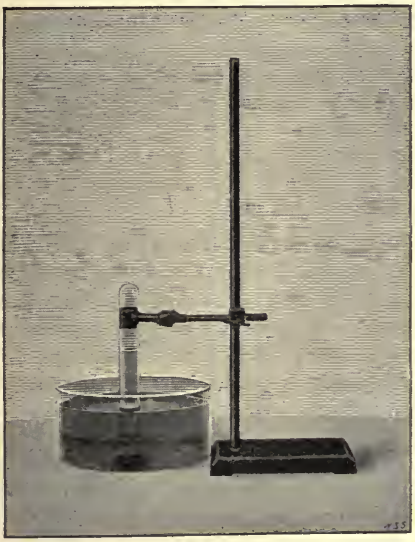

Fig. 36.

manner shown, and put the sodium, wrapped in the gauze, underneath. The hydrogen, as it is liberated, rises into the tube.° It will be found to burn with a yellow flame, the yellow colour being due to traces of sodium.

If potassium is thrown on water, a result similar to that pro-

duced by sodium is obtained, but so much heat is developed that the hydrogen takes fire. The flame is coloured violet by the potassium.

EXPT. 74. **To compare the weight of sodium hydrate with the weight of the sodium from which it is made.**

Weigh a small dry dish, put into it a clean piece of sodium about 0.2-0.4 grm., and weigh again as rapidly as possible. Take out the sodium, put a little water into the dish, and then throw the sodium on to the water, a very little at a time. When it has all disappeared, carefully evaporate off the water, dry the residue thoroughly, allow to cool and weigh. A decided increase in weight is observed.

It is possible by means of an electric current to re-obtain sodium from sodium hydrate, in fact sodium was originally prepared by this means.

The question might be asked, "Did the sodium expel all the hydrogen from the water on which it acted?" An answer to this question can easily be obtained.

EXPT. 75. **To show the liberation of hydrogen by the action of aluminium on sodium hydrate.**

Heat a little water with a few fragments of aluminium foil. No gas is evolved. Water is not affected by aluminium. Pour the water away, add a solution of sodium hydrate, and warm. A brisk effervescence is produced, and the gas given off is found to be hydrogen. Hence we see that every particle of water that is acted on by sodium to make sodium hydrate, loses only part of its hydrogen, and from the sodium hydrate that is made, more hydrogen may be obtained by the action of aluminium. It is found that as much hydrogen can be released from the sodium hydrate as was obtained by the action of sodium on water.

A method whereby both the hydrogen and oxygen may be obtained from water will be described later. (See p. 69.)

EQUATIONS.

(*a*) $Mg + H_2O = MgO + H_2$.
(*b*) $Fe_3 + 4H_2O = Fe_3O_4 + 4H_2$.
(*c*) $Na + H_2O = NaOH + H$.

QUESTIONS. X.

71. What is the substance formed by burning oxygen in hydrogen? By what tests or experiments could you establish the identity of the product?

72. Describe an experiment by which it may be proved that burnt hydrogen is water.

73. Describe two methods of obtaining hydrogen from water.

74. Why do we call oxygen and hydrogen chemical elements, and why do we consider water to be a compound of these two elements?

75. What is an oxy-hydrogen flame? What is the lime-light?

76. What difference is there between oxygen combined with hydrogen in water, and oxygen which is dissolved in water?

77. What is the effect of putting sodium on water? What is the appearance of sodium hydrate?

78. When a piece of sodium is put on water hydrogen is evolved. Why do you consider that the hydrogen comes from the water and not from the sodium?

CHAPTER XI.

COMPOSITION OF WATER.

I. By the Method of Cavendish.

CAVENDISH, who first proved that water was produced by the union of oxygen and hydrogen, showed also the proportions by volume in which the two components unite. He filled a bell-jar, standing over mercury and provided with a stop-cock, with a mixture of two volumes of hydrogen and one of oxygen. A pear-shaped vessel, made of thick glass, and through the top of which

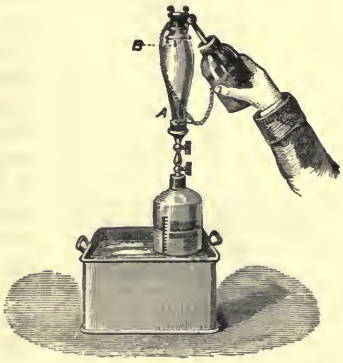

Fig. 37.

passed two platinum wires, was, at the narrow end, also provided with a stop-cock. After the air had been pumped out of this, it was connected with the bell-jar, and the taps opened, when the explosive mixture rushed in and filled it. The taps were then closed, and an electric spark was passed between the platinum wires. The oxygen and hydrogen united, and moisture was formed on the inside of the vessel. On connecting the two again, and opening the taps, more of the mixed gases entered the pear-shaped vessel, and was exploded as before. By proceeding in this way a considerable quantity of liquid was produced, which was proved to be pure water, and when the gases were made to combine in the above proportions, no gas, either hydrogen or oxygen, was left.

II. By the Eudiometer.

The hydrogen and oxygen may also be exploded in a **eudiometer**. This is merely a graduated tube, either straight or bent into a **U**-shape, made of thick glass, and sealed up at one end. Near the closed end two platinum wires are fused in the glass, so that the ends inside the eudiometer are nearly in contact. The eudiometer is first filled with mercury, and then the gases are introduced in turn. It is better not to have them in the proportion of two volumes of hydrogen to one of oxygen, as the explosion is then less likely to break the apparatus. If the **U**-shaped form is used, by bringing the mercury to the same level in both limbs before taking the reading, the gas may always be measured under atmospheric pressure. The open end is closed with a rubber stopper, held firmly in position, and an electric spark passed. A flash of light is

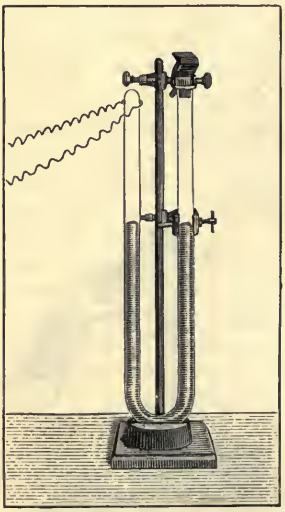

Fig. 38.

seen, but no loud report is heard. A small quantity of water is seen on the surface of the mercury after cooling. The mercury is again levelled, and the volume of residual gas is read. After filling the open limb with mercury, the residual gas may be transferred to the open end, and tested. It is found that the gas that has disappeared consisted of two volumes of hydrogen to one of oxygen.

The above shows the ratio of the volume of hydrogen to that of oxygen by "synthesis." In the following experiment the same point is proved by splitting the water up into its constituents, or by "analysis."

III. By Electrolysis.

When an electric current is passed through water acidulated with sulphuric acid—pure water does not allow the current to pass—it is decomposed, and both the oxygen and the hydrogen may be collected.

The apparatus shown consists of a **U**-tube, with the ends fitted with stop-cocks, and with a third tube ending in a bulb, joined on at the bend. Through each arm of the **U**-tube, and near the bend, a platinum wire, to the end of which a piece of platinum foil is welded, passes. The whole apparatus is called a **voltameter**, the platinum plates are the **electrodes**, and operation is called **electrolysis**.

The tubes are filled to the level of the taps with acidulated water, and the electrodes are joined to the terminals of an electric battery. The current travels from one pole of the battery along the wire to one

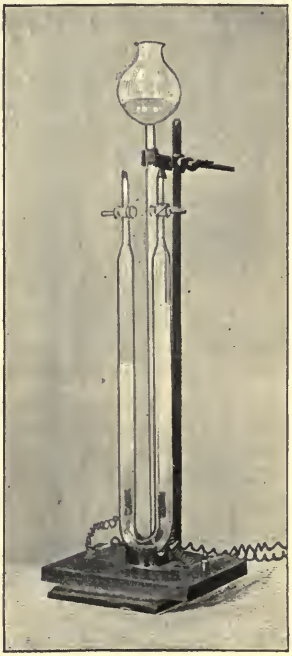

Fig. 39.

electrode—the **anode**—then through the liquid to the other electrode—the **kathode**—and so back to the other pole of the battery. It is found that the gases are liberated only at the electrodes, and as they collect in the closed tubes, it is soon seen that the volume in one tube is double that in the other. On testing, the lesser volume is proved to be oxygen, the other hydrogen.

IV. To show that two volumes of hydrogen unite with one volume of oxygen to form two volumes of steam.

So far nothing has been proved with regard to the volume of the water, in the form of steam, produced by the union of oxygen and hydrogen. When the gases are exploded in a eudiometer in the ordinary way, the steam condenses almost instantly, and forms a very small quantity of liquid water. But by keeping the tube at a temperature above 100° C., the volume of the steam may be compared with the volume of the constituent gases. This is done by surrounding the limb of the eudiometer, in which the explosion takes place, with the vapour of a liquid boiling above 100° C. Amyl alcohol, B.P. 132° C., is generally used.

The hydrogen and oxygen, in the proportion of two volumes of hydrogen to one of oxygen, are introduced into the closed limb which is then surrounded by a wider tube, through which the vapour for heating the gases passes. After passing the vapour for some time, the mercury is levelled in the two limbs, and the volume of mixed gases is read. The pressure is then diminished by letting some of the mercury run out, and the spark is passed. On levelling the mercury, it will be found that the volume of steam made is two-thirds of the volume of the mixed gases. That is to say, *two volumes of hydrogen unite with one volume of oxygen, and make two volumes of steam.*

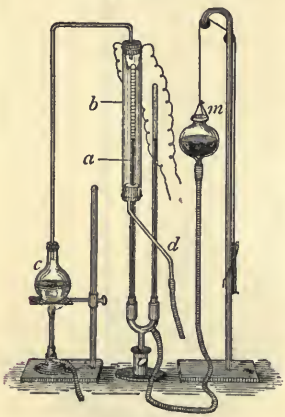

Fig. 40.

V. Composition of water by weight.

It is necessary also to find the relation between the *weights* of oxygen and hydrogen in water. This relation was first determined with accuracy by Dumas, a rough model of whose apparatus is shown in Fig. 41.

The hydrogen used was freed from all possible impurities, as far as could be done, special care being taken to remove all traces of moisture.

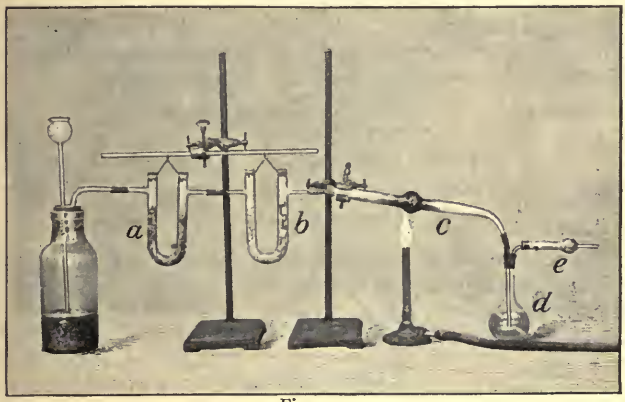

Fig. 41.

The purifying tubes are represented in the figure by the two U-tubes, *a* and *b*, *a* containing pumice soaked in silver nitrate solution, and *b* pumice soaked in strong sulphuric acid. The bulb-tube *c* contains dried copper oxide.

When hydrogen is passed over heated copper oxide, the oxygen is taken from the copper by the hydrogen, and water is formed. Most of this collects in the flask *d*, the remainder is caught in the small drying-tube *e*.

The copper-oxide tube is weighed before and after the experiment. It is found to have lost in weight, the loss representing the oxygen in the water formed. The flask and drying-tube *e* are weighed together also before and after the experiment. The

increase in weight represents the water formed. The hydrogen is then found by subtracting the loss in weight of *c* from the gain in weight of *d* and *e* together. The following figures refer to an experiment performed with the apparatus shown in the figure:

Wt. of flask and tube.

Second weighing = 47.449 grm. = wt. of flask + tube + water.
First „ = 44.942 grm. = wt. of flask + tube.
Difference = 2.507 grm. = wt. of water formed (*A*).

Wt. of copper oxide tube.

First weighing = 70.550 grm. = wt. of tube + copper oxide.
Second „ = 68.330 grm. = wt. of tube + copper.
Difference = 2.220 grm. = wt. of oxygen (*B*).

$$A - B = 0.287 \text{ grm.} = \text{wt. of hydrogen.}$$

Hence $$\frac{\text{wt. of oxygen}}{\text{wt. of hydrogen}} = \frac{2.220}{0.287} = \frac{7.74}{1}.$$

As a result of many experiments, Dumas came to the conclusion that oxygen and hydrogen unite in the proportions of 8 : 1 by weight, very nearly.

QUESTIONS. XI.

79. How did Dumas determine the composition of water by weight?

80. In an experiment by Dumas' method, 9.42 grams of water were produced. How much did the copper oxide lose in weight?

81. What weight of hydrogen is contained in 1 litre of water?

82. It was found in an experiment by Dumas' method that the copper oxide tube lost 11.421 grms. What weight of water would be produced?

83. Describe the method of determining the composition of water by the eudiometer.

84. Given the composition of water by weight and also by volume, calculate the relative densities of oxygen and hydrogen.

85. 20 ccs. of oxygen are mixed with 10 ccs. of hydrogen and exploded. What volume of gas, if any, will be left? How would you identify it?

86. What volume of air, containing 21% of oxygen, would be required to burn 5 litres of hydrogen?

87. 30 ccs. of air, containing 21% of oxygen, are mixed with 20 ccs. of hydrogen, and exploded. What volume of gas will remain, and of what will it consist?

88. 50 measures of air were mixed with 30 measures of hydrogen and fired by an electric spark. After cooling 48.8 measures of gas remained. What volume of oxygen was present in the volume of air examined?

89. Describe exactly any experiment *you have witnessed* illustrating the decomposition of a compound by means of the electric current.

CHAPTER XII.

OXIDATION AND REDUCTION.

THE changes that go on when iron rusts, when substances like phosphorus, a candle, sulphur, etc., burn in air or in oxygen, or when metals are heated in air, are all of the same kind. This fact has been long recognized, but up to the time of Lavoisier, the accepted explanation of these phenomena was wrong. It was said that when metals were calcined, and when various substances were burnt, something, which received the name of **Phlogiston,** was lost. Substances like hydrogen and carbon, which burn readily, were said to do so because they contained much phlogiston. To recover the original substance from the products of the burning, phlogiston had to be supplied to it.

When the balance came to be used, it was found that the products of burning were heavier than the original substances, and after that the true explanation of the changes was soon discovered. Lavoisier showed what part air plays in processes of burning (see p. 16).

We now know that the changes brought about are caused by the union with oxygen, or **oxidation,** the products being called **oxides.** The increase in weight represents the amount of oxygen with which the burning substance combines.

It is often stated that oxygen and air (on account of the oxygen in it) are "supporters of combustion," the substance burnt, *e.g.* hydrogen, or a candle, being described as "combustible bodies." Now the fact of the matter is, that when two substances are inclined to unite, it does not matter how they are presented to one another. Hydrogen burns at a jet in oxygen; mixed with oxygen the burning is over almost as soon as a light is applied to the mixture; and oxygen delivered at a jet burns quite readily in hydrogen. In each of these cases water, the product of the union of oxygen and hydrogen, is the only substance formed. Coal-gas burns in air at a jet, that is, the hydrogen and carbon, of which the coal-gas is principally composed, unite with oxygen from the surrounding air, forming water and carbon dioxide. If coal-gas is mixed with a proper amount of air and a light is applied, the union takes place almost instantaneously, the same substances being produced. We can also get air to burn in coal-gas.

To show air burning in coal-gas.

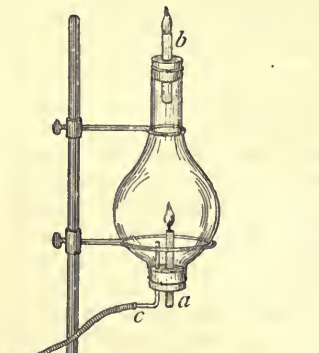

Fig. 42.

The lamp-glass, Fig. 42, is closed at the bottom by a cork, through which pass two tubes, one, *a*, wide and straight, the other, *c*, narrower and bent once at a right angle. The top is closed with a cork, through which a straight open tube *b* passes. This is at first closed, and coal-gas is led into the lamp-glass through the tube *c*. When it is full, the excess escapes through the tube *a*, at the end of which it can be lit. If then the tube *b* be opened, the flame is drawn up the tube *a*, on the upper end of which it remains. The excess of coal-gas now passes through the tube *b*, and may be burnt as it escapes.

We have then two flames, in one of which, viz. that in the lamp-glass, the air passing up the tube *a* plays the part of the "combustible body," the coal-gas, which surrounds it, acting as the "supporter of combustion." In the other we have the ordinary state of affairs, a jet of coal-gas burning in air. The action going on, viz. the union of hydrogen and carbon with oxygen, is the same in both cases. Had it happened that we required an atmosphere of hydrogen or coal-gas in which to live, then air or oxygen would be required to be delivered through the gas-pipes.

In nearly all cases union with oxygen is accompanied by the production of heat. In most cases this is very evident. But it often happens that oxidation goes on very slowly, as in the rusting of iron or the tarnishing of magnesium when exposed to the air, and then heat is produced at such a slow rate, that it is given off as fast as it is produced. Consequently no rise in temperature is noticed. But if the magnesium is heated, so as to induce rapid oxidation, the heat which results from the union is then evident. Similarly, if a handful of dead leaves be thrown on the fire, the hydrogen and carbon, of which they are chiefly composed, unite with oxygen to form water and carbon dioxide, and we are at once conscious that heat is being produced. If the same leaves were left exposed to the air at the ordinary temperature they would gradually decay, being oxidized to water and carbon dioxide. No rise in temperature would be noticed, because the heat would be dissipated as fast as it was produced. But the total amount of heat is the same whether the oxidation proceeds rapidly or slowly. The heat produced by decay is rendered sensible in hot-beds. When heaps of leaves or manure decay, the heat cannot all escape, and so the temperature is kept above that of surrounding objects.

The temperature of our bodies is kept above that of surrounding objects by the heat produced by oxidation. Food consists largely of compounds of carbon and hydrogen, and the body is chiefly composed of these substances. Oxygen, from the air we inhale in breathing, is taken up by the blood and carried to all parts of the system. Here it is used up in oxidizing various substances, carbon dioxide and water being the principal pro-

ducts. The carbon dioxide that is formed is carried in the blood back to the lungs, whence it is sent out into the air.

EXPT. 76. **To show that carbon dioxide is contained in the air we breathe out.**

Fit up the apparatus shown in Fig. 43. The test-tube contains a little lime-water, and is provided with a cork bored with two holes. A glass tube *a* passes nearly to the bottom of the test-tube, and another tube *b* passes just through the cork. Applying the mouth to the tube *b*, fill the lungs with air which will have to enter through the tube *a*, and bubble through the lime-water. No appreciable difference will be noticed (see, however, p. 127). Now apply the mouth to the tube *a*, and breathe out, again causing the air to pass through the lime-water. A milkiness is produced, and this we have learnt is a test for carbon dioxide.

Fig. 43.

Fishes require oxygen for breathing just as land animals do. They obtain it from the air which we proved to be dissolved in water in the experiment described on p. 26.

Reduction.

This is just the opposite to oxidation.

Substances that will take oxygen from bodies containing it are called "reducing agents," and the substance losing oxygen is said to be "reduced." Carbon and hydrogen are the reducing agents most generally employed; coal-gas, which is a mixture of free hydrogen with compounds containing carbon and hydrogen, is also a good reducing agent. When we passed

hydrogen over heated copper oxide in determining the composition of water, the hydrogen reduced the copper oxide, uniting with the oxygen to form water, and metallic copper was left. We were only concerned with the weight of oxygen and hydrogen in that experiment, and paid no attention to the weight of copper. By a slight modification we can determine the relative amounts of copper and of oxygen in black copper oxide.

EXPT. 77. **To determine the weight of oxygen combined with 1 grm. of copper in black copper oxide, by reduction.**

Weigh about 1 grm. of copper oxide in a small porcelain boat.

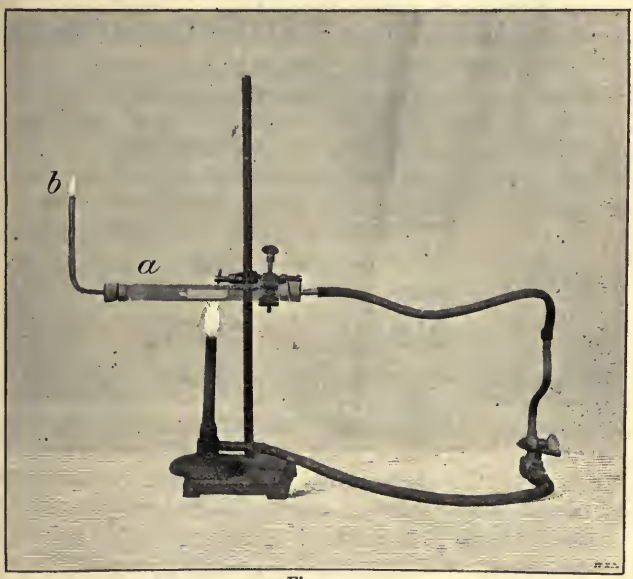

Fig. 44

Introduce this into a hard glass tube *a* (Fig. 44), through which hydrogen can be passed ; or ordinary coal-gas, which consists

chiefly of hydrogen and compounds of carbon and hydrogen, may be used. Allow the gas to pass until all the air is expelled, and then commence to heat the copper oxide. The rate at which the gas is passing should be regulated so that it gives a flame about 1 in. long when it is lit at *b*. The carbon and hydrogen in the coal-gas take the oxygen from the copper oxide, and in about ten minutes the reduction will be complete. Allow the tube to cool, letting the gas pass through it the whole time. When it is cold, take out the porcelain boat by means of a copper wire which has one end bent into a small hook, and weigh. Then heat for another five minutes in the coal-gas, cool, and weigh again. If this weighing differs by only a small amount from the previous one it may be regarded as correct.

Enter results in the following manner :

Second weighing	= 4.429 grm.	= wt. of boat + copper oxide.
First „	= 3.213 „	= wt. of boat.
Difference	= 1.216 grm.	= wt. of copper oxide.
Third weighing	= 4.185 grm.	= wt. of boat + copper.
First „	= 3.213 „	= wt. of boat.
Difference	= 0.972 grm.	= wt. of copper.

1.216 grm.	= wt. of copper oxide.
0.972 „	= wt. of copper.
0.244 grm.	= wt. of oxygen.

Hence weight of oxygen combined with 1 grm. of copper is

$$\frac{0.244 \times 1}{0.972} \text{ grm.} = 0.251 \text{ grm.}$$

EXPT. 78. **To determine the weight of oxygen combined with 1 grm. of copper in red copper oxide.**

Proceed exactly as in the last experiment. It will be found that the amount of oxygen combined with 1 grm. of copper is about 0.125 grm., *i.e.* half the amount combined with 1 grm. of copper in black copper oxide.

The results of the last two experiments illustrate an important

law of chemical combination which will be referred to later (see p. 173).

EXPT. 79. **To determine the weight of oxygen combined with 1 grm. of lead in yellow lead oxide (litharge).**

The reduction of litharge is easily carried out by means of coal-gas. It will be found that 1 grm. of lead combines with about 0.077 grm. of oxygen.

EXPT. 80. **To determine the weight of oxygen combined with 1 grm. of lead in lead dioxide (the 'puce-coloured" oxide).**

It will be found, employing the same method, that the weight of oxygen combined with 1 grm. of lead in this case, is double the amount in the last.

Most oxides may be reduced by means of hydrogen or coal-gas. Thus, if we pass hydrogen through a tube containing heated ferric oxide, or iron-rust, we find that water is produced, and iron is left in the form of a black powder. This iron is often in such a finely divided state that, on shaking it out into the air, it unites with the oxygen again so rapidly that it is made red-hot, and falls as a shower of sparks.

In the laboratory we often reduce a metallic oxide by the method employed in the following experiment.

EXPT. 81. **To obtain lead from litharge.**

Make a shallow depression in a piece of charcoal, and put in it a little litharge, or yellow lead oxide. Close the holes in the tube of the bunsen burner, and turn down the gas until a flame about 2 in. long is obtained. Put the nozzle of a mouth blow-pipe just at the edge of the flame, and blow gently (Fig. 45). The flame is diverted, and also altered in appearance. The inner part is yellow, the outer part is blue and non-luminous. By holding a thin piece of wire in different positions, it will be found that the flame extends some distance beyond the visible portion. Direct the flame so that the inner flame, which contains unburnt carbon, completely covers the litharge on the charcoal.

This is soon reduced, and bright shining globules are seen. Take off some of these, and give one a smart tap on the bench with the foot of the bunsen burner. It is flattened out. Rub it

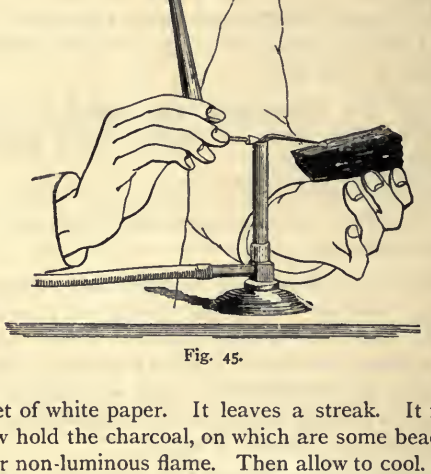

Fig. 45.

over a sheet of white paper. It leaves a streak. It is metallic lead. Now hold the charcoal, on which are some beads of lead, in the outer non-luminous flame. Then allow to cool. A yellow film of litharge is again seen. The inner blow-pipe flame is called the *reducing flame*; the outer flame, which allows air to have access to any substance heated in it, is the *oxidizing flame*.

EXPT. 82. **To obtain silver from silver nitrate.**

Heat a crystal of silver nitrate in the same way as the lead oxide in the last experiment. A shining globule of silver is obtained. It can be flattened out, but does not mark paper. Hold it in the oxidizing flame. It is not altered.

EXPT. 83. **To obtain antimony from antimony oxide.**

Heat a little oxide of antimony in the reducing flame. Shining globules are again obtained, consisting of metallic antimony. A smart blow causes these to break up, as arti-

mony is a very brittle metal. Hold a bead of antimony in the oxidizing flame. It is covered with a white film of antimony oxide.

QUESTIONS. XII.

90. What do you mean by "reduction"? Give examples of reducing agents.

91. What occurs when air is led through a tube containing heated copper? How would copper be again obtained from the product?

92. Describe with a sketch an arrangement for showing that air will burn in coal-gas.

93. Explain how the temperature of our bodies is kept above that of surrounding objects.

94. How would you obtain metallic lead from red lead? What are the principal properties of lead?

95. By what evidence can you show that one ingredient of the breath is identical with the gas formed by burning carbon in excess of air or oxygen?

96. Hydrogen is said to be a reducing agent. What does this mean? Mention some other reducing agents.

97. By what characters can lead, silver, antimony be distinguished from one another?

CHAPTER XIII.

THE FORMATION OF SALTS.

On burning substances in oxygen, we found that some of the products, or oxides, dissolved in water, and gave liquids which turned blue litmus red. These we called acids. Other oxides dissolved in water, and the liquid turned red litmus blue. These we called alkalies. We shall next try the effect of mixing an acid and an alkali together.

Expt. 84. **To prepare common salt by neutralizing hydrochloric acid with sodium hydrate.**

Half fill a test-tube with sodium hydrate solution, and then pour nearly the whole of the solution into a clean porcelain

dish. Half fill another tube with dilute hydrochloric acid solution, and then cautiously add it, a little at a time, to the sodium hydrate solution.

Stir with a glass rod after each addition, and after stirring, take out a drop of the solution on the end of the rod, and test the action on red litmus paper. At first the colour will change to blue, but presently it will be found that the red colour remains. The alkali now is apparently lost, and, tested by litmus, the liquid is acid. Now dip the stirring-rod into the solution of sodium hydrate left in the first tube, and transfer small quantities of it to the dish. Test with litmus paper after each addition. Very soon it will be found that the liquid in the dish is without action on either red or blue litmus.[a] The liquid is neither acid nor alkaline. Pour a small quantity into a test-tube, and add one drop of acid. The liquid obtained will turn blue litmus red. To another small quantity add one drop of the alkali. The liquid obtained in this case will turn red litmus blue. Pour a little of the liquid from the dish on to aluminium and warm. No action is produced. (We found when we were considering hydrogen, that either sodium hydrate or hydrochloric acid would, separately, dissolve aluminium, and give off hydrogen.) Boil the remainder of the liquid to dryness. A white solid is left. Taste it. It is recognized as ordinary table-salt, common salt, or sodium chloride.

EXPT. 85. **To neutralize hydrochloric acid with potassium hydrate.**

Start with potassium hydrate and hydrochloric acid, and proceed in exactly the same way.[b] A liquid can again be obtained which is without action on litmus, and which, on evaporating to dryness, leaves a white residue which has a distinctly "salty" taste.

In the above experiments, we say the acid and the alkali **neutralize** one another. In every case of this kind the product is called a **salt**. The salts produced when hydrochloric acid is neutralized are called chlorides. Thus we have made sodium chloride and potassium chloride in the above experiments. When nitric acid is the acid used, the salts produced are called nitrates, while sulphuric acid gives salts called sulphates.

To neutralize a definite amount of any alkali, a definite amount of acid is required. Had we taken twice as much sodium hydrate, we should have had to use double the amount of hydrochloric acid in Expt. 84. It may be asked: "Will equal amounts of different alkalies require the same amount of acid to neutralize them?" We shall test this in the case of sodium hydrate and potassium hydrate. For measuring the volumes of acid required, it will be convenient to use a burette.

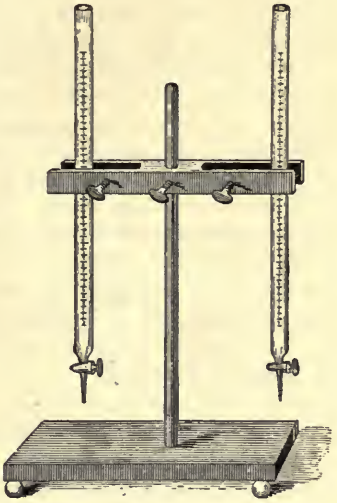

Fig. 46.

A **burette** is a piece of apparatus by which any required volume of a liquid may be measured. The form provided with a glass tap will be most convenient. It should be graduated to 50 c.cs., with each c.c. divided into 10 equal parts. Each small division will then represent 0.1 c.c. Before using the burette, wash it out with water, and then once or twice with a few c.cs. of the liquid to be used, letting the liquid run away through the tap. Then fix it vertically in the stand, and pour in the liquid. Take care that the narrow tube below the tap is full of liquid,

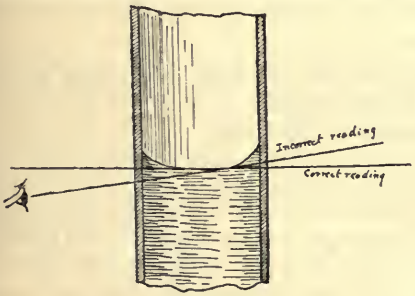

Fig. 47.

and that the upper surface of the liquid is in the graduated portion of the tube. When turning the tap with the right hand, steady the burette, by holding the part in which the tap fits, with the left. Take care not to pull the tap out.

As has been noticed when measuring the volume of a liquid in a measuring cylinder, the surface of water in a glass tube is curved. In order to get accurate results, the mark opposite the same part of the curve (usually the lowest point) must always be read, and the eye of the observer must be at the same height as the surface of the liquid. The error which is introduced by not attending to this last direction will be easily understood by reference to Fig. 47.

EXPT. 86. **To test whether equal volumes of a given solution of hydrochloric acid are required to neutralize 1 grm. of sodium hydrate and of potassium hydrate.**

Weigh a small flask, place in it about 0.5 grm. of solid potassium hydrate, and weigh again as rapidly as possible. Then dissolve the potassium hydrate in water, add a little litmus, and run in the solution of hydrochloric acid supplied from the burette till the alkali is neutralized. Note the volume of acid used, and calculate the volume that would be required to neutralize 1 grm. Proceed in exactly the same way, using sodium hydrate. It will be found that the volumes of acid required to neutralize 1 grm. of alkali are very different in the two experiments.

EXPT. 87. **To compare the volumes of a nitric acid solution required to neutralize 1 grm. of sodium hydrate and of potassium hydrate.**

Find the volumes of a given nitric acid solution required to neutralize 1 grm. of sodium hydrate[c] and 1 grm. of potassium hydrate.[d] It will be found that the *ratio* between the volumes is the same as in the last experiment, but the actual volumes themselves will probably be different. It would be the merest coincidence, if, in making the two solutions, the strengths were such that equal volumes were required.

When we burned sodium in oxygen, the sodium oxide tha

was made, dissolved in water, forming sodium hydrate. Potassium oxide also dissolves in water, forming potassium hydrate. It might be thought that the oxide of any metal would dissolve in water and give an alkaline liquid. We shall therefore test a few to see whether this is so or not.

EXPT. 88. Shake up a little calcium oxide with water. Allow the undissolved portion to subside, and test the clear solution with red litmus. It gives the reaction of alkalies.

EXPT. 88A. Shake up small quantities of black copper oxide, and yellow lead oxide (litharge) with water. If the undissolved portions do not settle readily, filter. Test a little of each filtrate with red litmus. No change is produced. Boil the remainder of both liquids to dryness. Practically no residue is left.

Most metallic oxides resemble those of lead and copper in being insoluble in water. Consequently, we are unable to test them with litmus. We can, however, try whether they possess the power of neutralizing acids.

EXPT. 89. **To prepare crystallized lead nitrate.**

Half fill a dish with dilute nitric acid and warm it. Add a small quantity of litharge, and stir with a glass rod until it dissolves.[e] Notice that no gas is given off. Continue adding litharge until some remains undissolved after stirring. Prepare a filter, using a funnel, the leg of which has been cut down to about $\frac{1}{4}$ in. Pour a little boiling water through this, and then filter the hot solution in the dish. If the filtrate does not deposit crystals on cooling, concentrate by evaporation, and again allow to cool. Pour off the liquid and dry the crystals with blotting-paper. Dissolve a few of the crystals in water, and test the solution with blue and red litmus paper. It will be found to be neutral. The white crystalline substance is a salt called lead nitrate.

EXPT. 90. Prepare copper sulphate crystals in precisely the same way, using black copper oxide and sulphuric acid.[f]

EXPT. 91. Prepare a specimen of zinc sulphate crystals, using zinc oxide and sulphuric acid.[g] Dissolve some of the

crystals in water, and test with red and blue litmus. The colour is not affected in either case.

The crystals made in the last experiment will be found to resemble, in appearance, those made by dissolving metallic zinc in sulphuric acid. We cannot at this stage make a complete analysis of the crystals, but we can at least determine the amount of water of crystallization.

EXPT. 92. **To test whether the crystals made by dissolving zinc in sulphuric acid are of the same composition as the crystals made by dissolving zinc oxide in sulphuric acid.**

Determine, as in Expt. 52, the amount of water of crystallization in 100 grms. of the crystals prepared in the last experiment. It will be found that the result is the same as was obtained on testing the crystals made by dissolving metallic zinc in sulphuric acid.

Other salts, e.g. *barium chloride*, *sodium sulphate*, *potassium nitrate*, *copper nitrate*, *might also be prepared.*

We see, then, that metallic oxides insoluble in water can neutralize acids as easily as the soluble oxides can. We call all those substances which possess this power **bases**, and we may state generally, therefore, that **bases neutralize acids, forming salts.** **Alkalies** are merely a small group of bases soluble in water.

This power of forming basic oxides is the most characteristic property of the elements we call **metals.** Other properties of the metals may be given here.

1. Metals are, with the exception of mercury, solid at the ordinary temperature.

2. They are opaque, and take a very high polish. We say they exhibit the "metallic lustre." Iodine and graphite, however, both non-metals, possess the metallic lustre. Gold, which is certainly a metal, in thin leaves is transparent.

3. Metals are, as a rule, heavy. Sodium and potassium, however, float on water, while lithium, a metal, is the lightest solid known.

4. Metals are the best conductors of heat and of electricity, although, in these properties, they differ widely among themselves.

5. Metals are, as a rule, malleable and ductile. It was seen in Expt. 83, however, that antimony was brittle.

It has been found that many metals dissolve in acids, and frequently hydrogen is given off. Now the metals, being elements, contain no hydrogen, and we must therefore conclude that the hydrogen comes from the acid. In making salts by neutralizing an acid with a base, that is, with the oxide of a metal, no gas was given off. What, then, becomes of the oxygen in the base, and of the hydrogen in the acid? We know that these two elements unite readily to form water, and, supposing they had done so in the preparation of the salts mentioned in this chapter, we should be able to understand why

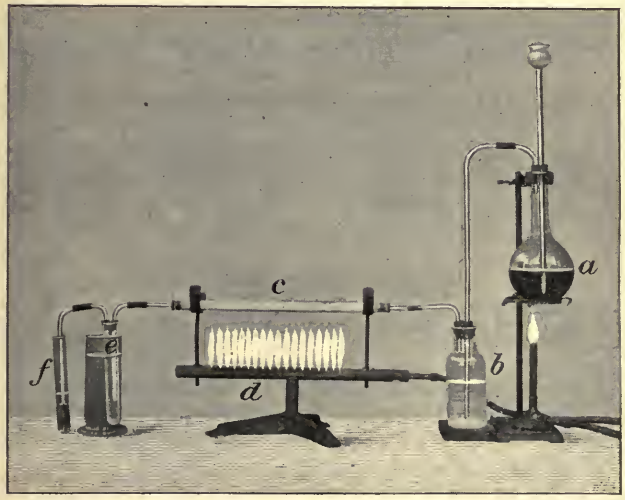

Fig. 48.

no gas was given off. We shall next describe an experiment which will enable us to prove whether this is so or not.

To demonstrate the formation of water when magnesium oxide neutralizes hydrochloric acid.[h]

We shall use hydrochloric acid, because this can be obtained in the form of a gas which is easily dried. In Fig. 48 the flask *a* contains rock-salt and sulphuric acid, which give hydrochloric acid, and this is dried by passing through sulphuric acid contained in the bottle *b*. The tube *c* contains an oxide, *e.g.* magnesium oxide, or lead oxide, which has been thoroughly dried by heating, and this tube can be heated gently by the burner *d* during the experiment. The tube *e* is placed in a vessel of cold water. As the hydrochloric acid passes over the oxide a salt is produced, which remains in the tube *c*. Water, which is at the same time produced, is carried forward, and condenses in the tube *e*. A little sodium hydrate solution is placed in the tube *f*, to absorb any acid which would otherwise escape into the air.

EQUATIONS.

(*a*) $NaOH + HCl = NaCl + H_2O$.
(*b*) $KOH + HCl = KCl + H_2O$.
(*c*) $NaOH + HNO_3 = NaNO_3 + H_2O$.
(*d*) $KOH + HNO_3 = KNO_3 + H_2O$.
(*e*) $PbO + 2HNO_3 = Pb(NO_3)_2 + H_2O$.
(*f*) $CuO + H_2SO_4 = CuSO_4 + H_2O$.
(*g*) $ZnO + H_2SO_4 = ZnSO_4 + H_2O$.
(*h*) $MgO + 2HCl = MgCl_2 + H_2O$.

QUESTIONS. XIII.

98. Iron rust and metallic iron are put into separate quantities of hydrochloric acid. How do the results in the two cases differ from one another?

99. Metallic copper when heated in air becomes coated with a black film. How would you prove that this black substance contains oxygen?

100. Describe two methods for the preparation of crystallised zinc sulphate.

101. Describe fully the method by which you have prepared any two crystallised salts.

102. Potassium nitrate, lead nitrate, lead chloride, and lead sulphate are described as "salts." What does this mean?

103. How would you prove that litharge contains lead and oxygen?

104. How would you prove that a mixture of black copper oxide and charcoal is not a chemical compound?

105. State the properties common to the metals, iron, lead, copper, and sodium, which distinguish them from such elements as carbon and sulphur.

106. Which of the following substances are metals: graphite, litharge, sulphur, copper, lime, iodine, iron, zinc, lead, mercury?

CHAPTER XIV.

CARBON.

To observe the effect of heating various compounds of carbon.

EXPT. 93. Carefully heat a little sugar in a dry tube. It is seen to melt and change colour, becoming yellow, brown, and finally black. The odour known as that of "burnt sugar" is noticed. Apply a light to the mouth of the tube. The gas given off takes fire, and burns with a yellow flame. The black mass left in the tube appears much more bulky than the sugar with which we started.

EXPT. 94. Heat a little flour, and a little sawdust, in the same way. The various results are practically the same as those in the last experiment.

Many other substances of vegetable or animal origin give similar results. When heated to a sufficiently high temperature they decompose, and a black residue is left. So-called "burnt" bread and "burnt" meat afford familiar examples of this action.

The black substance is called **carbon**. As obtained in this way, it is quite without definite shape. To indicate this, we use the word "amorphous," which means, literally, "without shape."

EXPT. 95. Heat small amounts of sugar, flour, sawdust, in iron spoons. Up to a certain point, the results are similar to those obtained on heating the same substances in test-tubes, but on continuing the heating the carbon disappears. It unites with oxygen from the air, forming carbon dioxide, as in Expt. 60.

More or less of a white or grey substance is left unaltered by heat. This is the ash.

The method of preparing various forms of amorphous carbon on a large scale is similar in principle to the above. Usually the material is heated in vessels made of iron or fire-clay, provided with an opening through which the gaseous products escape.

When coal is heated out of contact with air, the gas which is given off, after purifying, is ordinary "coal-gas." The residue is **coke**, which is largely used as fuel, and also in metallurgical operations. When heated strongly in air the carbon of the coke burns, and the ash is left.

EXPT. 96. Place in the bowl of a "churchwarden" pipe a few bits of coal, close the mouth of the pipe with clay, and heat the bowl strongly in a fire, or by means of a good bunsen burner. A gas will be given off, which may be lighted at the end of the stem, and after the operation is over, a residue of coke will be found in the bowl.

For preparing coal-gas, the coal is heated in long fire-clay retorts. It is found that a particularly dense form of amorphous carbon collects on the inside, towards the top of these retorts. This is called **gas-carbon**. It is used in voltaic batteries, and in making the arc light. When burnt it leaves only about 0.2 to 0.3 per cent. of ash.

Animal charcoal is the name given to the residue after bones are heated out of contact with air. It contains, besides carbon, the salts, the chief of which is calcium phosphate, to which bones owe their rigidity. When animal charcoal is heated strongly in air, the carbon burns away, and the salts are left as a white ash. Animal charcoal is used for removing colouring matter from liquids, *e.g.* in refining sugar. If a dilute solution of litmus be allowed to filter through animal charcoal

contained in a vertical tube, it will be noticed that the liquid runs out colourless.

Charcoal. When amorphous carbon is obtained from wood it is called charcoal. It is made in large quantities by firing stacks of wood (Fig. 49), which are almost completely covered with turf, so as to admit only a limited supply of air. Part

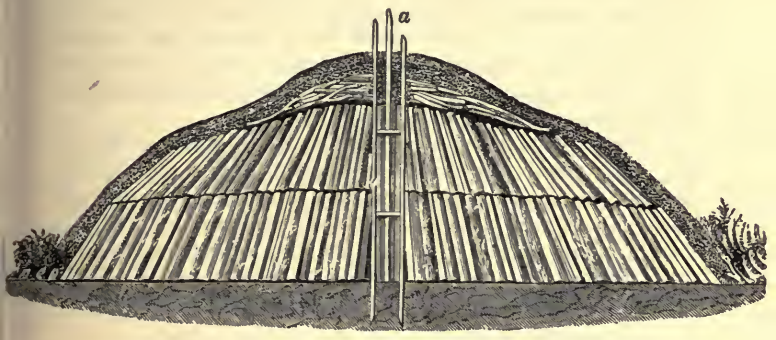

Fig. 49.

of the wood burns away, and the heat produced chars the remainder. On the Continent charcoal is largely used as fuel; it is also used in metallurgical operations. It has been found already that on burning charcoal in oxygen a white ash is left.

Soot consists largely of amorphous carbon. When specially prepared by burning substances very rich in carbon in a limited supply of air, it is called **lamp-black.** This substance is used in making printers' ink.

EXPT. 97. **To show that charcoal is heavier than water.**

Put a small piece of charcoal in a test-tube half full of water. It is apparently lighter than water. Press the charcoal under the surface by means of a second smaller tube, and boil the water for a few minutes. Bubbles of gas are given off long before the water boils. Cool the water, still keeping the charcoal underneath the surface. On removing the second tube

the charcoal is found to be heavier than water. In the first case the pores were filled with air, which buoyed up the charcoal. This air is removed when the charcoal is heated.

EXPT. 98. Powder some charcoal, and shake it in a vessel with water. On allowing the vessel to stand the charcoal sinks.

Charcoal is able, in some way or other, to absorb large quantities of different gases. On account of this property, charcoal is used as a disinfectant, the noxious gases being brought into intimate contact with the oxygen absorbed, and so rendered harmless. It is a striking fact that gases which are most easily liquefied are absorbed by charcoal to the greatest extent.

If a test-tube standing over mercury be filled with ammonia gas, and a piece of charcoal—it is best to re-heat the charcoal for this experiment—be introduced, the absorption of the ammonia will be marked by a rise of the mercury.

One volume of box-wood charcoal absorbs at 12° C. and 724 mm. pressure:

Ammonia,	90	vols.
Hydrochloric acid,	85	„
Sulphur dioxide,	65	„
Sulphuretted hydrogen,	55	„
Carbon monoxide,	9.42	„
Oxygen,	9.25	„
Nitrogen,	6.5	„
Hydrogen,	1.2	„

Wood charcoal is able to absorb colouring matter from liquids, although to a less extent than animal charcoal.

Graphite, known also as plumbago and black-lead, is a crystalline form of carbon. It is found chiefly in the United States, Russia, and Siberia. It is used for making black-lead pencils, for polishing iron-work, as a lubricant, and, mixed with fire-clay, for making plumbago crucibles.

Carbon dissolves in molten cast-iron, and in certain forms the carbon, on cooling, separates as graphite. On dissolving the iron the graphite is left.

Diamond is another crystalline form of carbon, although the crystals are of a different order from those of graphite. It is

the hardest substance known. The clear colourless diamonds, when cut and polished, are gems. Black diamonds, which are as hard as the colourless variety, are used on the ends of cutting and boring tools, and also for grinding and polishing the colourless ones. Diamonds are found in India, South Africa, and Brazil.

Small diamonds have been made by rapidly cooling iron in which carbon is dissolved. The outer layer becomes solid first, and contracts, and so the carbon in the interior crystallizes out under great pressure.

The different forms of carbon are proved to be of the same composition by burning equal weights in oxygen, when it is found that the same weight of the same substance, carbon dioxide, is produced in each case. The production of carbon dioxide by burning charcoal in oxygen has already been proved.

It is found that the greater the specific gravity of the carbon, the higher is the temperature at which it will begin to burn. The specific gravities of some forms of carbon are given below:

	Sp. Gr.
Diamond, - - - - - -	3.514
Gas-carbon, - - - - -	2.356
Graphite, - - - - - -	2.11 to 2.26
Charcoal, - - - - - -	1.45 to 1.7

At the ordinary temperature carbon is very inert, not uniting with any other element. At high temperatures it unites directly with a number of substances, *e.g.* hydrogen, oxygen, sulphur. It will also take oxygen from oxides, and in this way it is applied in the extraction of many metals from their oxides.

Carbon occurs in the substance of all animals and plants. Compounds of carbon are so numerous, that they are, with the exception of the oxides and the carbonates, considered in a separate branch of chemistry, viz. Organic Chemistry.

QUESTIONS. XIV.

107. By what experiments can it be shown that charcoal, graphite, and diamond consist essentially of the same material?

108. What is the composition of the substances called "red lead" and "black lead" respectively? What would be the effect of strongly heating a mixture of red lead and black lead?

109. A splinter of a colourless transparent solid body is given to you. How would you ascertain (*a*) by physical tests, (*b*) by chemical tests, whether this is glass or diamond?

110. Some sawdust is heated to redness in a test-tube. Describe all that is seen during the course of this experiment.

111. State the natural sources of graphite, and describe those properties which distinguish it in a marked degree from amorphous carbon, and which render it especially useful. It what arts and manufactures is it usefully employed?

112. Give some account of the different varieties of amorphous carbon, and mention the uses to which they are applied.

CHAPTER XV.

CARBON DIOXIDE AND CARBONATES.

WE found that by burning carbon in oxygen a gas was produced which turned lime-water milky, and which, when dissolved in water, gave a liquid that turned blue litmus red. This gas, which we called carbon dioxide, is produced when any substance containing carbon is burnt in air or in oxygen.

EXPT. 99. **To prove that carbon dioxide is made by burning compounds containing carbon in air.**

Introduce a lighted candle, fixed on the cup of a deflagrating spoon, into a dry jar containing air. The candle soon goes out. Moisture is seen on the sides of the jar, and if the residual gas is shaken up with lime-water, the presence of carbon dioxide is shown. Why does the candle not continue to burn? We know that, in burning, the substance of which the candle is composed is uniting with oxygen, and so a very natural answer is that the oxygen has been used up, and carbon dioxide formed.

We shall try and find whether this is the correct answer later (see Expt. 119).

Expt. 100. Hold a lighted splint in a jar of air. When the light goes out, the lime-water test shows that carbon dioxide has been made.

Expt. 101. Hold a jar for a short time over a small gas-flame. The presence of carbon dioxide is shown as before.

Carbon dioxide is usually obtained from compounds called carbonates.

Expt. 102. **To show that many carbonates give off carbon dioxide on being heated.**

Heat a small quantity of chalk, or powdered marble, in a small hard glass test-tube.[a] Hold inside the tube a glass rod, on the end of which is a drop of clear lime-water. This becomes turbid, showing the presence of carbon dioxide. In the same way test magnesium carbonate, white lead, zinc carbonate, sodium carbonate, barium carbonate, or sodium bicarbonate. It is found that most carbonates give carbon dioxide when heated, while a small number will not.

Expt. 103. **To show the formation of carbon dioxide by the action of an acid on a carbonate.**

Put a little sodium carbonate into a test-tube, and pour on it hydrochloric acid.[b] A brisk effervescence is set up, and the liberation of carbon dioxide is indicated as before. The gas may, instead, be poured into a test-tube containing lime-water, just as water would be, and then shaken up with the liquid.

Try also the action of sulphuric acid, nitric acid, and acetic acid, respectively, on various carbonates. Dissolve also a little tartaric acid, and try this solution in the same way. In each case carbon dioxide is evolved.

To prepare the gas in quantity, marble (calcium carbonate) is acted on by hydrochloric acid.[c]

Expt. 104. **To fill several jars with carbon dioxide.**

Set up an apparatus similar to that shown in Fig. 50.

Let the pieces of marble, each about the size of a nut, slide down the side of the bottle, and cover them with water. Pour down the thistle funnel, the end of which must be under the surface of the water, commercial hydrochloric acid. The delivery tube should pass through a hole in a piece of cardboard, which rests on the mouth of the jar to be filled. This method of collecting a gas is the method of "downward displacement." Collect several jars of the gas, covering the mouth of each jar with a glass plate. What property of carbon dioxide may be inferred from the mode of collecting it?

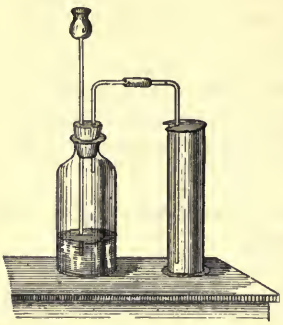

Fig. 50.

EXPT. 105. **To show the action of carbon dioxide on a lighted candle.**

Introduce a lighted candle, fixed on the cup of a deflagrating-spoon, into the gas. The gas does not take fire, neither does it allow the candle to burn.

EXPT. 107. **To show that carbon dioxide is heavier than air.**

Pour the gas from one jar to another exactly as you would pour water. Prove the presence of the gas in the lower jar by means of lime-water.

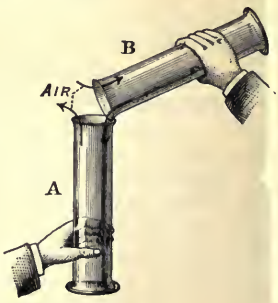

Fig. 51.

EXPT. 107. Pour the gas downwards on to a lighted splint. The light is extinguished.

By means of a lighted taper we may show the gradual rise of carbon dioxide in a jar in which it is being collected.

EXPT. 108. While a jar is being filled, introduce a lighted

taper at short intervals. At first, it will be found that the taper is not extinguished until near the bottom of the jar, but afterwards it goes out nearer and nearer the mouth.

That carbon dioxide is heavier than air may be shown in other ways. Thus if a bell-jar be filled with carbon dioxide (the gas should first be washed by bubbling through water), soap-bubbles blown with air will float on it.

If a beaker, placed mouth upwards on one arm of a balance, be counterpoised, and carbon dioxide be then poured into it, an

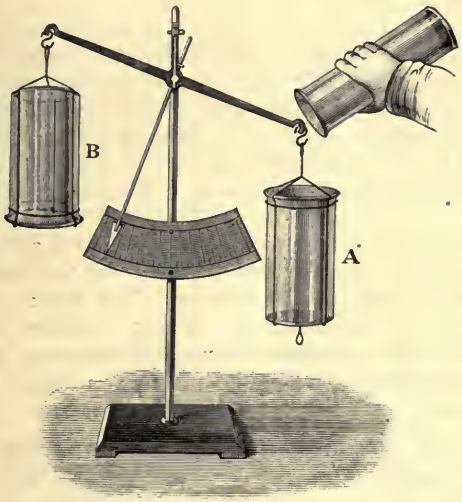

Fig. 52.

increase in weight is at once evident. If the beaker is left for abo t half-an-hour, it will be found that equilibrium is restored, the carbon dioxide having escaped.

EXPT. 109. **To show that carbon dioxide slowly makes its way out of an open vessel placed mouth upwards.**

Fill five or six dry test-tubes with carbon dioxide and leave

them in the stand, mouth upwards, and with the mouths uncovered. Test them in turn at intervals of five minutes. It will be found that the carbon dioxide slowly makes its way out.

Carbon dioxide is about 1½ times as heavy as air, and can be poured downwards like water, yet the last experiment shows that we cannot keep the gas in an uncovered jar, even if the mouth is turned upwards. The gas slowly makes its way upwards and out into the surrounding air. Compare the corresponding experiment with hydrogen, p. 42.

EXPT. 110. Allow carbon dioxide to bubble through water. As we already know, from Expts. 60 and 65, it dissolves, and the solution turns lime-water milky, and reddens blue litmus. The solution is said to contain an acid called carbonic acid. Warm some of the solution in water. Bubbles of gas escape, and are proved to be carbon dioxide by the lime-water test. Also boil the solution coloured with litmus. The colour changes to purple.

At the ordinary temperature water will dissolve about its own volume of carbon dioxide, and if it is collected over water, therefore, there is always considerable loss. The amount dissolved is increased by increasing the pressure. Carbon dioxide is contained under a pressure of several atmospheres in various effervescing drinks. When the pressure is released carbon dioxide escapes, and this causes the effervescence.

EXPT. 111. **Examination of the residue left in the carbon dioxide apparatus.** Filter the solution left in the bottle in which the gas was generated, evaporate to a small bulk, and allow it to cool. Colourless transparent crystals separate. Pour off the liquid, and dry the crystals with blotting-paper. They do not resemble marble in the least. Dissolve part of the crystals in water, and test parts of the solution with blue and red litmus. No change is produced. Heat a few crystals in a dry tube. Water of crystallization is given off. The residue is white and amorphous. When it is cold add a drop or two of water. A considerable rise in temperature is noticed.

The substance obtained in this experiment is called **calcium chloride.** The residue left after the water of crystallization is removed from calcium chloride crystals readily unites with water. It is often used for removing moisture from gases. We used it to dry the hydrogen in Expt. 70.

Carbonates.

Seeing that carbon dioxide is an acid forming oxide, we should expect it to unite with any alkali, and form a salt. We shall therefore try its action on sodium hydrate.

EXPT. 112. Select a test-tube of such a size that it can easily be covered by the thumb, and fill it with carbon dioxide. Pour in a little sodium hydrate, immediately cover the mouth of the tube tightly with the thumb, and shake. The carbon dioxide will be absorbed, as is shown by the thumb being sucked in.[d]

The acid and the alkali unite with one another, and produce a salt called sodium carbonate. Potassium carbonate might be made in the same way.

Now add hydrochloric acid to the solution. The effervescence which is set up is easily proved to be due to the escape of carbon dioxide.

EXPT. 113. Pass carbon dioxide into a considerable volume of lime water.[e] Allow the precipitate to settle, pour off the liquid, and to the solid residue add hydrochloric acid. A brisk effervescence is set up. Pour the gas evolved into lime-water. It is turned milky. The white precipitate is evidently a carbonate, insoluble in water.

All carbonates are decomposed by acids, carbon dioxide being evolved. Most carbonates are decomposed on being heated, carbon dioxide being again evolved. Of the common carbonates, only those of sodium, potassium, and ammonium are soluble in water. If a solution of any one of these be added to a solution containing a salt of another metal, we get a precipitate consisting of the carbonate of this metal.

EXPT. 114. Add sodium carbonate solution to a solution of calcium chloride.[f] Wash the precipitate that is produced, and

add to it a little hydrochloric acid. Test the gas which is evolved with lime-water. It proves to be carbon dioxide; we have, therefore, by mixing the two solutions, made a carbonate which is insoluble in water. This has been brought about by the metals changing places,

sodium carbonate + calcium chloride *yielding* sodium chloride + calcium carbonate (insoluble).

EXPT. 115. Add a solution of sodium carbonate to a solution of copper sulphate. A green precipitate is obtained. Wash this precipitate and add to it dilute hydrochloric acid. Pour the gas which is evolved into lime-water. It is turned milky.

Hard and Soft Water.

EXPT. 116. **To show that calcium carbonate is dissolved by water containing carbon dioxide.**

Pass carbon dioxide into lime-water as in Expt. 113, and continue the passage after the white precipitate is formed. The liquid gradually becomes clear again. Boil part of this clear solution in a test-tube. Test for carbon dioxide in the escaping steam by means of lime-water on a glass rod. The lime-water becomes turbid. A white precipitate also re-appears in the tube, and if tested with hydrochloric acid, proves to be a carbonate. Calcium carbonate, therefore, although insoluble in ordinary water, is soluble in water containing carbon dioxide,[g] but when, by any means, this carbon dioxide is removed, the calcium carbonate is again thrown down.

EXPT. 117. To another part of the solution prepared in the last experiment, add more lime-water. A white precipitate is thrown down. The lime reacts with the carbon dioxide in solution, forming calcium carbonate.[h] This is precipitated, and along with it there is also the calcium carbonate originally formed.

The solution of calcium carbonate is carried on on a large scale in nature. When organic substances decay, carbon dioxide is one of the products, and so when water passes over the surface of the earth and through the earth, it takes carbon

dioxide into solution. If then, this water comes in contact with chalk or lime-stone, some of the solid dissolves, to be re-precipitated when the carbon dioxide is removed, either by boiling or by other means. When calcium carbonate is deposited on the

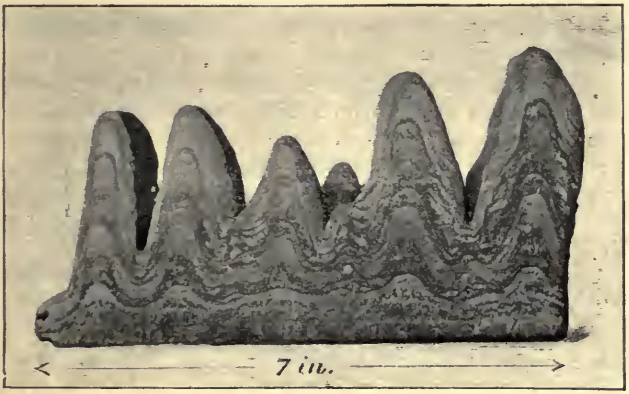

Fig. 53.—Specimen of boiler-crust.

interior of steam boilers, it constitutes "boiler-crust," and if left to accumulate for a long time, may attain a great thickness (Fig. 53).

When water containing calcium carbonate percolates through the roof of a cave, as each drop hangs for a short time, a little carbon dioxide is given off, and a trace of the solid is deposited. This going on for a long time, causes rods of calcium carbonate, called "stalactites," to be formed, which hang from the roof like icicles. Similar formations can often be seen under bridges. When the drops of water fall to the ground, more carbon dioxide is given off, and "stalagmites" form on the floor.

The action of a "petrifying well" is of the same kind; objects over which the water flows are gradually covered with a deposit of calcium carbonate.

Water which contains calcium carbonate in solution is distinguished as "hard" water, that is, it does not easily give a lather with soap. Before it does lather, a white curd, which

floats on the surface of the water, is formed, and this represents a certain waste of soap. The calcium carbonate is precipitated either by boiling or by adding more lime (Expt. 117), and thus the water may be softened. The hardness is therefore temporary.

To compare the hardness of different samples of water.

Different samples of water are compared with respect to hardness, by finding what volumes of the same soap solution are required to give a lather with equal volumes of the samples.

For practice in making these comparisons, hard water may be prepared by passing carbon dioxide through ordinary water to which different amounts of lime-water have been added, until the precipitate is re-dissolved, or, more easily, by dissolving known weights of calcium chloride in known volumes of water.

A soap solution may be prepared by dissolving, say, 10 grm. of Castile soap in a litre of weak methylated spirit.

EXPT. 118. Measure 50 c.cs. of distilled water, by means of a measuring cylinder, into a stoppered bottle. Run in the soap solution from a burette, a little at a time, and shake after each addition until a lather remains unbroken on the surface after the bottle has stood for a minute. The volume of soap solution used represents the amount needed to give a lather with 50 c.cs. of water free from calcium carbonate. Test the samples supplied in precisely the same manner, running in the soap solution 1 c.c. at a time until the end is nearly reached. Then add the soap solution most cautiously until a lather similar to that obtained with the distilled water is formed. Deduct from the volume of solution required, the amount needed to give a lather with the distilled water.

Suppose 1 c.c. of soap solution gave a lather with the distilled water, and that two other samples required respectively 9 c.cs. and 13 c.cs. The hardness of these two samples will be as 8 : 12.

EXPT. 119. **To find the least percentage of carbon dioxide in the air necessary to prevent a candle burning.**

Measure the capacity of a gas-jar by finding the volume of water it will hold. Then test mixtures of air and carbon

dioxide in different proportions, with a lighted candle. We know the candle goes out when the jar is full of carbon dioxide. Try then, at first, a mixture containing 50 per cent. of carbon dioxide. Half fill the jar with water—the remainder will be air—and displace the water by carbon dioxide over the pneumatic trough. Remove the jar and introduce the lighted candle. It is at once extinguished. Try next a mixture of 25 per cent. carbon dioxide and 75 per cent. of air. This also extinguishes the candle at once. Continue to try mixtures with smaller percentages of carbon dioxide until you find one which does not extinguish the candle at once.

Then try with larger amounts of carbon dioxide until you find the mixture which just does extinguish the candle as soon as it is introduced. The various percentages should each be tried twice. Of course, the result of an experiment such as this would be expected to depend in an unusual degree on the experimenter, but it will be found that very few get results different from 6 to 8 per cent. of carbon dioxide. Of the remaining gas, we know that about one-fifth is oxygen, so that we see it is far from the truth to say that a candle goes out in a closed jar because *all* the oxygen is used up.

Care should be taken to fill the jar with water before every experiment, so as to make sure that all the carbon dioxide from the previous experiment has been removed.

EQUATIONS.

(*a*) $CaCO_3 = CaO + CO_2$.

(*b*) $Na_2CO_3 + 2HCl = 2NaCl + H_2O + CO_2$.

(*c*) $CaCO_3 + 2HCl = CaCl_2 + H_2O + CO_2$.

(*d*) $CO_2 + 2NaOH = Na_2CO_3 + H_2O$.

(*e*) $Ca(OH)_2 + CO_2 = CaCO_3 + H_2O$.

(*f*) $CaCl_2 + Na_2CO_3 = CaCO_3 + 2NaCl$.

(*g*) $CaCO_3 + H_2O + CO_2 = CaCO_3H_2CO_3$ (soluble).

(*h*) $CaCO_3H_2CO_3 + Ca(OH)_2 = 2CaCO_3 + 2H_2O$.

QUESTIONS. XV.

113. How could you show that the gas contained in human breath, which renders lime-water turbid, is also contained in limestone, marble, oyster-shells?

114. Write a summary of the properties of carbon dioxide.

115. Describe the method you employed in filling a jar with carbon dioxide.

116. If you place a burning candle in a closed bottle, the flame will soon be extinguished. Why does the candle not continue to burn? How is the composition of the air affected by the flame?

117. A mixture of chalk and sand is given to you. Describe and explain a method whereby these two substances could be separated from one another.

118. Describe experiments with carbon dioxide to show (*a*) its density, (*b*) its action on lime-water, (*c*) its solubility in water.

119. What is the meaning of the term "hard" as applied to water? Explain how it comes about that rain, which is nearly pure water, usually becomes hard after falling on the earth.

120. In what respects do hard, soft, mineral, and distilled waters differ from one another? In what class do you place Thames water? How would you distinguish distilled water from a mineral water?

121. How could you show that when charcoal burns in air it is the oxygen and not the nitrogen that sustains the combustion?

122. How would you prove that carbon is present in (1) sugar, (2) carbon dioxide, (3) calcium carbonate?

123. Describe the properties of any four carbonates.

124. How would you compare the hardness of any two samples of water?

125. How do you account for the "furring" that occurs in kettles when certain kinds of water are boiled in them?

CHAPTER XVI.

CARBON MONOXIDE.

Composition of the Oxides of Carbon.

UNDER certain circumstances, carbon dioxide may be deprived of the whole or part of its oxygen. Thus, if carbon dioxide, dried by bubbling through strong sulphuric acid, be passed through a tube in which a little potassium is being heated, the latter burns and a black residue containing carbon is left. On shaking this in water, an alkaline solution is produced.

To reduce carbon dioxide to carbon monoxide by means of carbon.

Carbon itself is able to take part of the oxygen from carbon dioxide. An iron tube *a*, of $\frac{5}{8}$ in. bore, is filled with fragments of charcoal and heated in a furnace. Carbon dioxide generated from marble and hydrochloric acid in the bottle *b*, is passed through the tube, and the issuing gas, after passing through a

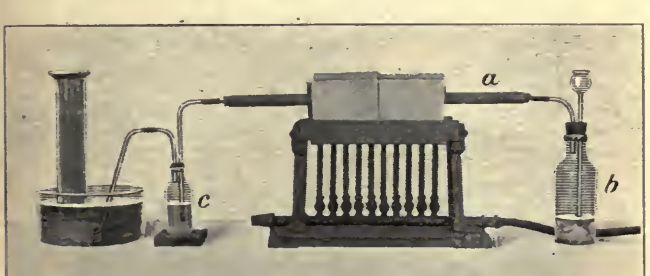

Fig. 54.

wash-bottle *c*, containing sodium hydrate solution, to remove the unaltered carbon dioxide, is collected over water. On shaking the gas with lime-water, no precipitate is produced. On applying a light, the gas burns with a blue flame, and on again shaking, the lime-water becomes turbid.[a]

This new gas is called **carbon monoxide**, and the blue flame produced by burning carbon monoxide is often seen on the top of a bright fire. The air combining with the fuel in the lower part of the grate forms carbon dioxide, which, passing through the fire, comes in contact with more red-hot carbon, and so carbon monoxide results. Charcoal fires are very liable to produce carbon monoxide.

Carbon monoxide is more conveniently prepared in the laboratory by the action of strong sulphuric acid on formic acid or oxalic acid, or on salts of these acids.

EXPT. 120. Add strong sulphuric acid to a little sodium formate in a test-tube.[b] Test the gas evolved with a drop of

lime-water on the end of a glass rod. No turbidity is produced. Apply a light to the mouth of the tube. The blue flame of carbon monoxide is seen.

Carbon monoxide is a highly poisonous gas, and care must be taken not to breathe it.

EXPT. 121. Heat a little oxalic acid with strong sulphuric acid, and test as before.[e] Carbon dioxide is readily detected in this case, but the gas, nevertheless, will burn with a blue flame. We have here, therefore, a mixture of carbon monoxide and carbon dioxide.

Before starting the next experiment it will be well to practise transferring gas from one tube to another. Half fill a test-tube with water, and invert it, mouth downwards, under water in the trough. Bring a

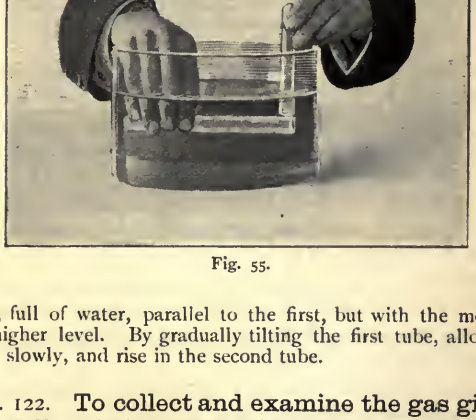

Fig. 55.

test-tube, full of water, parallel to the first, but with the mouth at a slightly higher level. By gradually tilting the first tube, allow the air to escape slowly, and rise in the second tube.

EXPT. 122. **To collect and examine the gas given off when oxalic acid is heated with strong sulphuric acid.**

Fit up an apparatus as shown in Fig. 56. A six-ounce flask is a convenient size. Put in the flask 8 or 10 grms. of oxalic acid, and cover with strong sulphuric acid. Take the

biggest tube that you can cover tightly with your thumb, and find another whose capacity is rather less than half that of the first. Warm the flask, and when all the air has been expelled, fill the larger of the two tubes with the gas. Have ready a tube in which 1 or 2 c.cs. of sodium hydrate solution

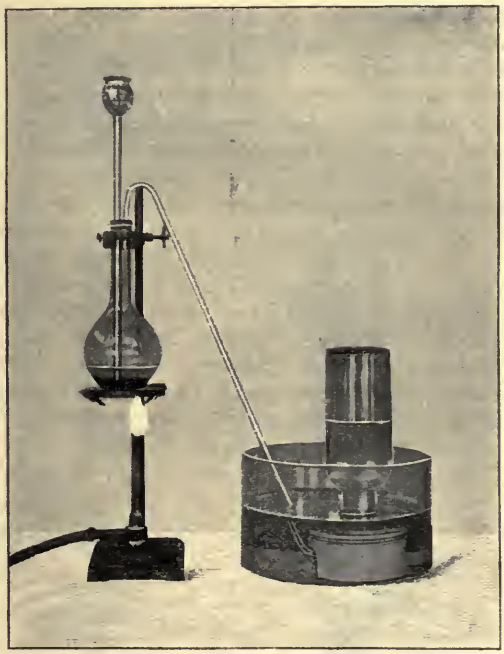

Fig. 56.

have been put. Pour the sodium hydrate as quickly as possible into the tube containing the gas, *at once* cover the mouth of the tube, and shake vigorously. The absorption of the carbon dioxide is shown by the thumb being sucked in. Open the tube under water. The water will be found to rise practically half-way up the tube. Transfer the remaining gas to the

smaller test-tube, pour in a little lime-water, and shake. No turbidity is produced. Apply a light to the mouth of the tube, and after the flame has died out shake again.[d] The lime-water is turned milky.

The above experiment shows most of the properties of carbon monoxide. It is a colourless gas, insoluble in water. It is not absorbed by sodium hydrate, nor does it turn lime-water milky. It burns with a blue flame, producing carbon dioxide. The gas from oxalic acid consists of equal volumes of carbon monoxide and carbon dioxide. When collected over water an appreciable amount of the carbon dioxide is dissolved, but if some of the gas is collected over mercury, the sodium hydrate is found to remove just one half.

If steam is passed through a red-hot tube containing charcoal, the carbon takes oxygen from the water, and an inflammable gas, consisting of a mixture of carbon monoxide and hydrogen, is produced.[e] This mixture constitutes "water-gas."

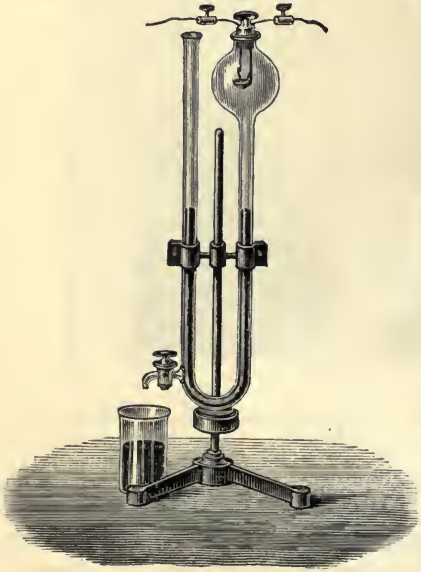

Fig. 57.

To show the relation between the volume of carbon dioxide, and the oxygen which it contains, the apparatus shown in Fig. 57 may be used. Mercury is poured in until it nearly reaches the bottom of the large bulb. Through the stopper two copper wires pass, to the bottom of one of which

a small cup is fixed. On this a little charcoal, as pure as possible, is put, and a thin platinum wire, connecting the two copper wires, is placed in contact with it. The stopper is put in position, and mercury is run out through the tap to diminish the pressure inside the bulb. The stopper must be quite gas-tight. On joining the two copper wires to the terminals of a small battery the platinum is made red-hot, and this fires the charcoal, which unites with the oxygen producing carbon dioxide. After cooling down it will be found that the volume of gas is the same as before. *Carbon dioxide contains its own volume of oxygen.*

The places where the copper wires pass through the stopper may be made gas-tight by a little melted india-rubber.

To show the relation between the volume of carbon monoxide, and of the oxygen which it contains, the gas is exploded with oxygen in a eudiometer. From 10 to 15 c.cs. of carbon monoxide and about an equal volume of oxygen are introduced into a eudiometer and exploded, as in the case of the mixture of hydrogen and oxygen. The volume of the residual gas is measured after cooling. The open limb is then filled with mercury, and a small piece of solid sodium hydrate is put on the top. The sodium hydrate is then transferred to the closed limb, where it absorbs the carbon dioxide that has been produced. On again levelling the mercury in the two limbs the extent to which the volume has been diminished can be read off. This diminution represents the volume of carbon dioxide produced. It will be found to be equal to the volume of the original carbon monoxide. We know also from the last experiment that it contains its own volume of oxygen. The gas still left will be found to be oxygen, and if the volume of this be deducted from the original volume of oxygen, we obtain the volume with which the carbon monoxide has combined. This last volume is found to be equal to half the volume of the carbon monoxide.

The following figures refer to an actual experiment made in this way:

Before Explosion.

Volume of carbon monoxide, - - -	11.3 c.cs.
„ oxygen, - - - - -	10.0 c.cs.

After Explosion.

Volume of residual gas, - - - - -	15.6 c.cs.
„ gas left after absorbing carbon dioxide with sodium hydrate, - - - - -	4.3 c.cs.
Difference = carbon dioxide produced,	= 11.3 c.cs.

The 4.3 c.cs. of gas proved to be oxygen.

$\therefore$ 10.0 − 4.3 = 5.7 c.cs. of oxygen had been used.

Hence we see that carbon monoxide unites with half its volume of oxygen, and produces its own volume of carbon dioxide. This latter we know to contain its own volume of oxygen, and therefore the carbon monoxide must have contained half its volume of oxygen. In equal volumes of carbon monoxide and carbon dioxide we have, therefore, combined with the same amount of carbon, amounts of oxygen which are as 1 : 2.

EQUATIONS.

(*a*) $CO_2 + C = 2CO$.

(*b*) $CH_2O_2 + H_2SO_4 = H_2SO_4 + H_2O + CO$.

(*c*) $H_2C_2O_4 + H_2SO_4 = H_2SO_4 + H_2O + CO + CO_2$.

(In the last two reactions the sulphuric acid determines the removal of the elements of water.)

(*d*) $CO + O = CO_2$.

(*e*) $C + H_2O = CO + H_2$.

QUESTIONS. XVI.

126. Write a summary of the properties of carbon monoxide.

127. Two bottles are given to you, one of which contains carbon monoxide, and the other carbon dioxide. By what tests or experiments could you identify the gases?

128. How would you distinguish carbon monoxide from hydrogen?

129. Air is passed through a tube full of common charcoal and heated to redness. How would you find out whether the air is changed in volume or in properties? Would it make any difference if in one experiment you use a very long tube of charcoal, and in another a very short tube?

130. Describe what happens on heating oxalic acid with strong sulphuric acid. What is the composition of the gas evolved?

131. How is carbon dioxide converted into carbon monoxide, and carbon monoxide into carbon dioxide?

132. How do you account for the blue flame which is often seen at the top of a coke fire?

133. How would you show that carbon dioxide contains its own volume of oxygen?

134. What substances are produced when charcoal is heated to redness with (*a*) steam, (*b*) carbon dioxide, (*c*) litharge?

135. How may pure carbon monoxide be obtained? What are the characters by which it may be recognized?

136. How would you prove that the gas obtained by heating oxalic acid with strong sulphuric acid contains equal volumes of carbon monoxide and of carbon dioxide?

137. 10 c.cs. of carbon monoxide are exploded with 12 c.cs. of oxygen. Of what will the residual gas consist? How would you determine its composition?

138. What occurs when carbon monoxide is passed over heated copper oxide?

CHAPTER XVII.

QUANTITATIVE EXPERIMENTS WITH CARBONATES.

EXPT. 123. **To find the weight of carbon dioxide given off when a carbonate is heated.**

Weigh in a small crucible, provided with a lid, about 1 grm. of powdered marble. As a very high temperature is required to liberate all the carbon dioxide,[a] the crucible, without lid, is conveniently heated in a small muffle furnace. The lid should be replaced as soon as the crucible is taken out, and the final weighing made as quickly as possible after the crucible is cold. Calculate the loss in weight on 100 grms. It is found that 100 grms. of marble always give off 44 grms. of carbon dioxide.

EXPT. 124. To the substance left in the crucible after heating marble, add a few drops of water.[b] It becomes very hot, and after cooling again, notice that a *dry* white powder is still left.

Limestone has the same composition as chalk or marble, and enormous quantities of limestone are roasted, or "burnt," as it is termed, in manufacturing **quick-lime.** This, as we have just seen, unites with water in the production of great heat, while the resultant substance still appears quite dry. This is **slaked lime.**

EXPT. 125. Shake the slaked lime with half a test-tubeful of water. It is evidently not very soluble in water. Allow the undissolved portion to settle, and pour off the clear liquid. This is **lime-water.** Add a little red litmus. It becomes blue. Also shake a little chalk in water, allow it to stand, and then pour off the clear liquid. Divide this into two parts, and test with red and blue litmus. Neither is affected. On heating, therefore, the neutral substance chalk gives off carbon dioxide, and leaves a white solid, quick-lime, which unites energetically with water, and gives a substance, slaked lime, which is soluble in water, forming an alkaline solution.

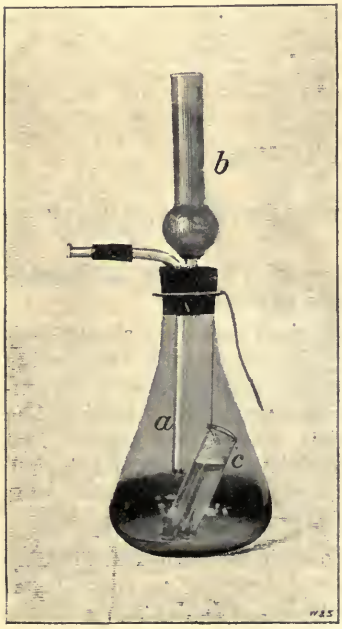

Fig. 58.

The determination of the loss in weight when magnesite, zinc carbonate (which is a basic salt), or white lead is heated, will afford additional practice. A bunsen flame may be used for heating these substances.

EXPT. 126. **To find the weight of the carbon dioxide given off when a carbonate is decomposed by an acid.**

For this experiment, the apparatus shown in Fig. 58 may be used.

The conical flask of 2 oz. capacity, is provided with a cork with two holes. Through one passes a glass tube *a*, reaching nearly to the bottom of the flask. The upper end of this tube is closed by a piece of rubber tubing and a glass stop. Through the other passes a little tube *b*, containing cotton-wool or glass-wool. A small test-tube *c* hangs by a piece of thread, which passes between the stopper and the neck of the flask.

Cover the bottom of the flask with water, put a few c.cs. of nitric acid in the small tube, and weigh. Then remove the stopper, taking care that none of the acid escapes from the small tube. Put in the flask about 1 grm. of powdered marble, replace the stopper, and weigh again. Now tilt the apparatus slightly, until a little nitric acid runs out of the tube. The acid acts on the marble, and the carbon dioxide as it escapes passes through the cotton wool, which retains the spray caused by the effervescence. Continue adding the acid until all the marble is decomposed.

The apparatus, when first weighed, was filled with air, now it is full of carbon dioxide. Suck this out through the narrow glass tube, taking care that the end of this is *not* below the surface of the acid. Then weigh again. The result will be entered as follows:

Second weighing = 38.41 grm. = wt. of apparatus + marble.
First „ = 37.34 grm. = wt. of apparatus.
Difference = 1.07 grm. = wt. of marble.

Second weighing = 38.41 grm. = wt. of apparatus + marble.
Third „ = 37.94 grm. = wt. of apparatus + (marble − carbon dioxide).
Difference = 0.47 grm. = wt. of carbon dioxide.

∴ Percentage of carbon dioxide in marble is

$$\frac{0.47 \times 100}{1.07} = 43.9.$$

It will be found that the weight of carbon dioxide given off in this experiment agrees with that found in Expt. 123.

Expt. 127. **To find the volume of carbon dioxide obtainable from a carbonate,** the apparatus shown in

Fig. 59 may be used. The bottle *a* is provided with a rubber stopper, through a hole in which passes a short piece of glass tube, bent once at a right angle. A similar bottle *b* is provided with a stopper with two holes. Through one passes a glass tube

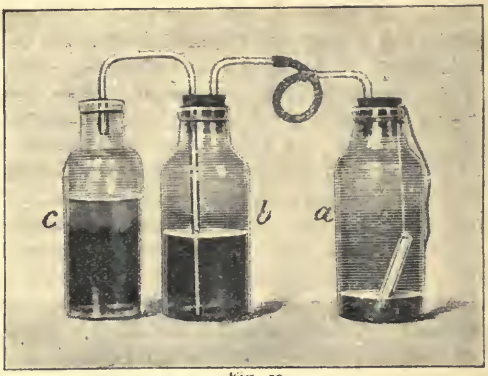

Fig. 59.

similar to that in the bottle *a*. These tubes are joined by rubber tubing. The second tube through the stopper passes to the bottom of the bottle *b*.

Weigh about 1 grm. of the carbonate into the bottle *a*, and cover it with water. Put a few c.cs. of strong hydrochloric acid in a small test-tube, and hang this by a thread passing between the stopper and the neck of the bottle. Fill the bottle *b* with water, fix the stoppers firmly, *and put the exit-tube from b in position only when everything else is ready.* By cautiously tilting the bottle *a*, allow a little acid to run out of the test-tube. Gas will pass out into the bottle *b*, expelling its own volume of water, which is caught in a third bottle *c*. Continue to add the acid till all the carbonate is decomposed, shaking the bottle *a* towards the end of the operation. The gas passing out of the bottle *a* will be air mixed with a little carbon dioxide, but a volume of carbon dioxide equal to the air expelled will remain in the bottle *a*. Consequently, if the volume of water in the bottle *c* be measured, we shall know the volume of carbon dioxide produced,

Powdered chalk, sodium carbonate, and sodium bicarbonate are suitable substances for this determination. Calculate the volume obtained from 1 grm.

In Expt. 113 it was seen that the white precipitate formed when carbon dioxide was passed into lime-water, effervesced with an acid, giving off carbon dioxide. The precipitate obtained in the lime-water test for carbon dioxide is therefore a carbonate, and, from the mode of its formation, it is probably calcium carbonate. But in Expt. 123 it was found that calcium carbonate lost 44 per cent. by weight when heated. Try, therefore, whether the white solid, obtained by the action of carbon dioxide on lime-water, loses the same amount.

EXPT. 128. **To show that calcium carbonate is produced by the action of carbon dioxide on lime-water.**

Pass carbon dioxide through a considerable volume, say 50 c.cs., of lime-water. When a thick precipitate has formed, filter it off, wash thoroughly, and dry. Weigh some of the dry precipitate in a porcelain crucible, and heat in the muffle. It will be found to have lost about 44 per cent. by weight.

We have learnt, therefore, that on heating calcium carbonate, say marble or limestone, carbon dioxide is given off, and a white solid, quick-lime, is left; that quick-lime unites with water with great energy, giving a white powder, slaked lime; that slaked lime dissolves in water, giving an alkaline liquid, called lime-water; and that when carbon dioxide is passed into this liquid, white insoluble calcium carbonate is again produced.

EQUATIONS.

(*a*) $CaCO_3 = CaO + CO_2$.
(*b*) $CaO + OH_2 = Ca(OH)_2$.
(*c*) $Ca(OH)_2 + CO_2 = CaCO_3 + H_2O$.

QUESTIONS. XVII.

139. Describe and explain the process of lime-burning.

140. How is calcic hydrate produced? What compound is produced by exposing the hydrate to the action of carbon dioxide?

141. What is the effect of adding water to (1) limestone, (2) quicklime?

142. Sketch and describe an apparatus for determining the percentage of carbon dioxide in any carbonate.

143. What is the effect of heating (1) sodium carbonate, (2) sodium bicarbonate, (3) calcic carbonate?

144. What occurs when lime is slaked? How do you account for the heat which is produced?

CHAPTER XVIII.

OTHER COMPOUNDS OF CARBON.

EXPT. 129. **To obtain alcohol by fermentation.**

Into a flask of 16 oz. capacity, put a solution of 20 grms. of sugar in 100 c.cs. of luke-warm water, and add a little yeast. Fit a cork, and a bent tube passing into a vessel containing lime-water. Put the whole away till next lesson. It will probably be noticed that the liquid in the flask has frothed up somewhat. If the lime-water is not turbid, boil it. A white precipitate will appear which has been formed by the action of carbon dioxide on the lime-water and afterwards re-dissolved when a further quantity of the gas was produced. Pour the liquid from the flask into another flask, leaving as much as possible of the yeast

Fig. 60.

behind. Throw this yeast away, and pour the liquid backwards and forwards between the two flasks several times. Then put it into a retort, and distil until about 12 or 15 c.cs. of distillate have been collected. Try if a little of this will burn. If it will not, add a little solid potassium carbonate to the distillate in a test-tube, shake well, and allow to stand. Two layers of liquid will form, and if the top layer be tested it will be found to burn. It consists of alcohol, or "spirits of wine."

The alcohol in wine is produced by the fermentation of a sugar contained in the grape juice; that in beer from a sugar made from the starch contained in barley.

Under the microscope, yeast is seen to consist of small round bodies called yeast-cells. These grow by a process of budding, and during their growth certain sugars—cane sugar is not one—are changed into alcohol and carbon dioxide. Ordinary yeast contains a substance which has the power of changing cane sugar into other kinds,[a] which, by the action of yeast, are split up into alcohol and carbon dioxide.[b] Alcohol boils at a lower temperature than water, and so, on distilling, the first part of the distillate will contain most of the alcohol. Potassium carbonate dissolves in water, but not in alcohol, and the latter, being lighter than the solution of the carbonate, floats on the top. This top layer will not be entirely free from water. To obtain dry alcohol, the alcohol obtained in the experiments of one or more classes may be collected and put into a flask half-filled with pieces of quick-lime. After standing for a day or two, the alcohol is distilled from a water-bath. The alcohol will be found to boil at 78° C.

EXPT. 130. To test whether alcohol is free from water or not, heat a few crystals of copper sulphate until all the water of crystallization has been given off. The residue is white. Add a little water to a small quantity of this powder. The blue colour reappears. Now add a little alcohol to another quantity. If the alcohol is free from water the colour will remain unchanged.

EXPT. 131. Light a few drops of alcohol, contained in the cup of a deflagrating-spoon, and place it in a dry gas-jar.

Moisture is deposited on the sides. Shake up with lime-water This is made turbid. Alcohol, then, contains carbon and hydrogen. It also contains some oxygen, whose presence cannot easily be demonstrated.

Alcohol dissolves many substances, *e.g.* shellac and iodine, which are insoluble in water. A few examples of its solvent action have already been noticed (see p. 21). Alcohol solidifies at an exceedingly low temperature, and is used in making low-temperature thermometers. It is also used as a preservative.

Methylated Spirit is spirits of wine to which about 10 per cent. of another liquid, wood spirit, has been added, together with a little naphtha. These are added to render the spirit unfit to drink, and do not interfere with its use in the many commercial operations in which it is employed. The naphtha causes a turbidity when water is added. Methylated spirit is sold free of duty, whereas a heavy duty is levied on alcohol used for drinking purposes.

Acetic Acid. If a fermented liquor, such as beer, is left exposed to the air in a shallow vessel, a mould gradually forms on the surface, and the liquid becomes sour. If tested with litmus, it gives an acid reaction. The change is brought about by the growth of a microscopic body called the "vinegar plant." These bodies are found floating about in the air, and when any fall into suitable alcoholic liquors, such as beer or light wines, they multiply, and during their growth the alcohol is changed into an acid called acetic acid. The liquid obtained when dilute alcoholic liquors turn sour under the influence of the vinegar-plant is vinegar.

Acetic acid is a colourless liquid, with a very pungent odour. Its boiling point is 118° C. The strongest acid, called glacial acetic acid, forms an ice-like mass when cooled to 16.5° C. It produces blisters when brought in contact with the skin. The preservative action of vinegar is due to the acetic acid which it contains.

Acetic acid possesses all the properties of an acid. It neutralizes bases, forming salts called acetates. From these the acid can be prepared by distilling with sulphuric acid.

EXPT. 132. Heat about 20 grms. of crystallized sodium acetate on an iron tray. The salt melts, and much water of crystallization is given off. Keep the substance well stirred, and when it is dry put it into a 6 oz. retort, cover it with strong sulphuric acid, and heat. Acetic acid distils over, and condenses as a colourless liquid, possessing a pungent odour.[e]

EXPT. 133. Pour a few drops of acetic acid into half a test-tubeful of water and taste the solution. It tastes like vinegar.

EXPT. 134. To a little of the distillate add blue litmus. It is turned red.

EXPT. 135. Pour a little of the distillate on a few fragments of zinc. A brisk effervescence is set up, and the gas given off takes fire. It is hydrogen.

EXPT. 136. Put the remainder into a porcelain dish, and add lead oxide as long as it will dissolve. Filter while hot. On allowing the filtrate to cool, white crystals of lead acetate, or "sugar of lead," separate.

EXPT. 137. **Methane.** Dry some sodium acetate as before,

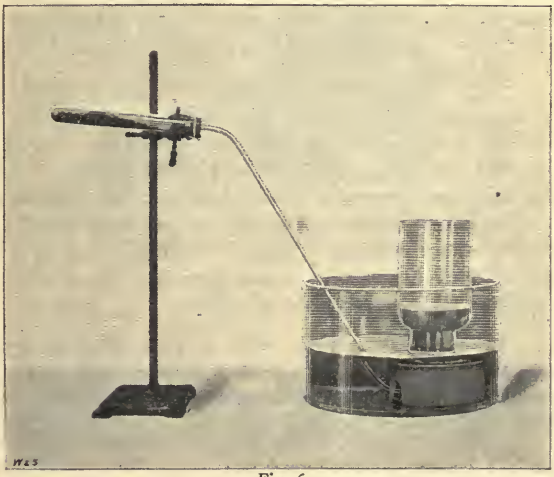

Fig 61.

and mix it with about twice its weight of soda-lime.[1] Put the mixture into a hard glass test-tube provided with a cork and delivery tube, and heat. Have the mouth of the test-tube turned slightly downwards. A colourless gas is given off, which can be collected over water.[d] It is called methane. Fill a few test-tubes with the gas, keeping them mouth downwards.

EXPT. 138. Shake the gas with lime-water. It is not turned milky.

EXPT. 139. Apply a light to the mouth of one of the tubes. The gas burns with a blue non-luminous flame.[e] Add a little lime-water and shake. A turbidity is produced.

EXPT. 140. Pour the gas upwards into a dry test-tube and apply a light. Evidently the gas is lighter than air, and it will be noticed that moisture is formed inside the tube when the gas is burnt. Methane contains carbon and hydrogen, and nothing else.

Methane is formed when vegetable substances decay, and may be seen rising in canals or stagnant pools. It is often called "marsh-gas." It constitutes the "fire-damp" of coal-mines. Methane is one of the constituents of the inflammable gas produced when coal or wood is heated out of contact with air.

EXPT. 141. **Ethylene.** To about 15 c.cs. of methylated spirit, cautiously add six times the volume of strong sulphuric acid, cooling well during the operation. Pour the mixture into a 10 oz. flask provided with cork and delivery tube as shown. On heating this mixture, a colourless gas called **ethylene** is given off, which is collected over water in the usual way.[f] Two or three test-tubes of the gas may be collected. Before all the air has escaped a mixture of ethylene and air is collected. This explodes on the application of a light.

EXPT. 142. Shake the gas with a little sodium hydrate solution, then transfer it to another test-tube, and shake with lime-water. No turbidity is produced. Apply a light. Notice the difference between the flame it produces and that of burning

[1] Soda-lime is made by slaking quick-lime with sodium hydrate solution and heating to drive off the excess of water.

methane. Shake with the lime-water again. The lime-water becomes milky.

(*N.B.*—The preliminary shaking with sodium hydrate is necessary, because changes occur in this preparation which

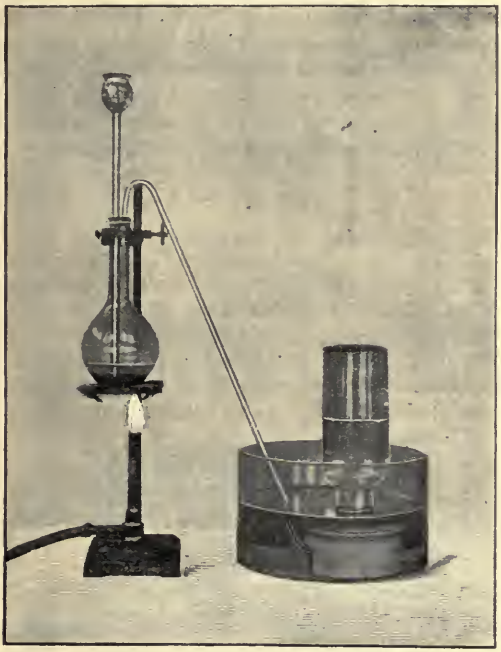

Fig. 62.

result in the formation, among other things, of a little carbon dioxide.)

Ethylene is present to a slight extent in coal-gas, and the luminosity of a coal-gas flame depends partly on the presence of ethylene.

Ethylene is composed of carbon and hydrogen, but it is found that the amount of carbon combined with a certain amount of

has also been weighed. The bulbs *A* contain a strong solution of potassium hydrate, the U-tubes, *B* and *C*, contain pumice soaked in strong sulphuric acid. The tube *ab* is heated in a furnace, and the stop-cocks are then slightly opened. Air passes into the apparatus, being first freed from carbon dioxide in the bulbs *A*, and then dried in passing through the tubes *B* and *C*, after which it passes through the tube *ab*. Here the copper unites with the oxygen, and the nitrogen passes into the globe *V*. At the conclusion of the experiment the apparatus is allowed to cool, and then the globe, and the tube *ab*, are weighed. The increase in weight of the globe gives the weight of nitrogen, and the increase in weight of *ab* gives the weight of oxygen. The proportions of the two gases by weight are found to be,

Oxygen, 23
Nitrogen, 77
100

Preparation of nitrogen from the air.

To prepare nitrogen, we may pass air over heated copper as in the experiment just described, and collect the nitrogen by displacement over water, or we may remove the oxygen by means of phosphorus, as in Expt 13. A much more rapid method is the following :

Fig. 64.

EXPT. 143. Cover the bee-hive shelf, standing in the pneumatic trough, to the depth of about ½ in. with water. Put a piece of dry phosphorus, rather larger than a pea, into a porcelain crucible standing on the bee-hive shelf. Kindle the phosphorus by touching it with a hot wire, and at once place an inverted gas-jar over it. When the phosphorus ceases to burn,

shake the jar about so as to dissolve the phosphorus pentoxide. The crucible will be filled with water at the same time, and sink. It will be seen that the water has risen in the jar, owing to the removal of the oxygen by its combining with the phosphorus. Transfer some of the residual gas to a test-tube in the same way that the carbon monoxide in Expt. 122 was transferred from one tube to another. Add a little lime-water to this and shake. No milkiness is produced. Then take the jar containing the remainder of the gas out of the trough and introduce a lighted splint. It is at once extinguished.

The gas in the jar is nitrogen. It is colourless, almost insoluble in water, and has neither taste nor smell. Substances which burn in air are extinguished when placed in nitrogen. Animals die almost immediately when put in a jar of nitrogen. As nearly four-fifths of the air we breathe consists of this gas, it cannot in itself be poisonous. Animals are killed when put into nitrogen, not by the nitrogen, but because of the absence of oxygen. The function of the nitrogen in air appears to be to dilute the oxygen.

The oxygen and nitrogen in the air are mixed, not actually combined. This is proved in many ways, of which three may at this stage be appreciated.

1. *The ratio of nitrogen to oxygen is not always the same;* samples of air from different localities are found to exhibit differences in composition, though only slight.

It has been already seen from numerous examples, and all experience goes to show, that when two substances unite to form a third, the ratio of the amounts of the two uniting substances is invariable.

2. *Air can be liquefied by cooling to a very low temperature. As the temperature rises, the liquid evaporates, but the nitrogen evaporates more rapidly than the oxygen, so that the last quantity of gas given off is able to rekindle a glowing splint.*

When a compound of two substances, *e.g.* water, is boiled, the gas has just the same composition as the liquid from which it is formed.

3. *When air is dissolved in cold water, and afterwards driven out by boiling, as in the Expt. described on p.* 26, *it is*

has also been weighed. The bulbs *A* contain a strong solution of potassium hydrate, the U-tubes, *B* and *C*, contain pumice soaked in strong sulphuric acid. The tube *ab* is heated in a furnace, and the stop-cocks are then slightly opened. Air passes into the apparatus, being first freed from carbon dioxide in the bulbs *A*, and then dried in passing through the tubes *B* and *C*, after which it passes through the tube *ab*. Here the copper unites with the oxygen, and the nitrogen passes into the globe *V*. At the conclusion of the experiment the apparatus is allowed to cool, and then the globe, and the tube *ab*, are weighed. The increase in weight of the globe gives the weight of nitrogen, and the increase in weight of *ab* gives the weight of oxygen. The proportions of the two gases by weight are found to be,

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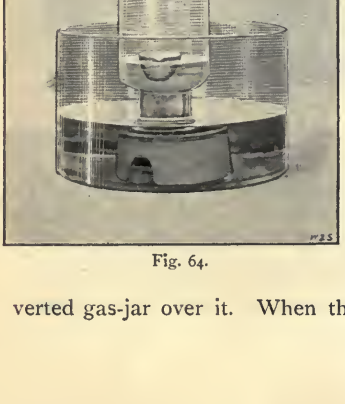

Fig. 64.

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3. *When air is dissolved in cold water, and afterwards driven out by boiling, as in the Expt. described on p.* 26, *it is*

found that the boiled-out air has not the same composition as the original air.

Oxygen is, under the same conditions of temperature and pressure, more soluble in water than nitrogen, and consequently the air dissolved in water contains a larger proportion of oxygen than is present in the original air. The amount is increased from 23 to 35 per cent. Nothing of this kind is observed when a compound gas is dissolved in water. The carbon dioxide, for example, given off when a solution in water is heated, has just the same composition as the original gas.

To show that air boiled out from water contains more oxygen than ordinary air.

The proportion of oxygen in air dissolved in water may be shown in the following way. A narrow glass tube, closed at one end, is filled with the gas boiled out from water as in the Expt. described on p. 26. A few c.cs. of a strong solution of pyrogallol are rapidly poured into the tube, and then a small piece of caustic potash is dropped in, and the tube is tightly closed with a rubber stopper. After shaking for about 10 min. the tube is opened under water. Alkaline pyrogallol rapidly absorbs oxygen, and the water rises in the tube, therefore, to take the place of the oxygen. The calculation of the percentage of oxygen in the sample of gas used is made in the same way as in Expt. 13.

By absorbing the oxygen with alkaline pyrogallol instead of with phosphorus, the determination of the amount of oxygen in ordinary air can be made very rapidly. (See Expt. 13.)

Besides the oxygen and nitrogen, however, other gases occur in the atmosphere to a greater or less extent. When carbon, and substances containing carbon, burn, carbon dioxide is produced. The burning of coal and coal-gas causes large quantities of carbon dioxide to be continually sent into the air. Much of the carbon in food, also, is changed ultimately to carbon dioxide in the bodies of animals, and this is exhaled by them. (See Expt. 76.) Consequently we might expect to find carbon dioxide in the air.

EXPT. 144. **To show the presence of carbon dioxide in the air.**

Expose a little lime-water to the air in a shallow glass dish for an hour or so. A white scum forms on the surface. This we have previously used as a test for carbon dioxide.

We learn now, therefore, that the reason why, in Expt. 76, the lime-water was not appreciably altered in appearance when we drew air through it, was not because the air contained no carbon dioxide, but because the amount in the volume of air we inhaled was so small. The average amount of carbon dioxide in the air is less than four parts in 10,000, or less than ·04 per cent.

It might be thought, since oxygen is being continually used to form carbon dioxide, that the air would be rendered unfit for breathing. Such indeed would be the case if there were no counteracting influence. But plants exert on the air an effect which is just the opposite to that of animals. Plants take all the carbon which they require for building up their structure from the carbon dioxide found in the air. They breathe through small openings in their leaves, and then the green substance which occurs in plants, by some means or other, decomposes the carbon dioxide, the carbon being retained, while the oxygen is evolved.

The action of plants on the atmosphere is such an accurate balance to that of animals and the effects of combustion, that the amount of carbon dioxide in the air is surprisingly constant. In large towns the amount of carbon dioxide in the air is naturally greater than in the country, and in crowded rooms the amount may be increased very largely. As we are continually using up oxygen in our bodies, it is very necessary that fresh air should be continually admitted to a room, particularly where large numbers of people are collected. The various systems of ventilation are intended to do this as far as possible without producing draughts.

Water-vapour in the air. If water is left exposed to the air in a shallow vessel, it gradually evaporates. Large quantities of water from the sea, rivers, and lakes continually disappear in this way. Hence it is not surprising to find that water, in the form of gas, is always present in the air to a greater or less

extent. This is shown to be the case in many ways. The inside of the windows of a warm room are often covered with moisture in cold weather. The outside of a glass of cold water brought into a warm room soon becomes covered with water. These effects show (1) that there is moisture in the air, and (2) that the air cooled by contact with the cold glass cannot hold as much moisture as warm air.

EXPT. 145. Weigh a watch-glass on which a little fused calcium chloride has been placed, and after leaving it exposed to the air for an hour or so weigh it again. A decided increase is observed. It will be seen that the surface of the calcium chloride appears wet, and if left long enough, sufficient water will be absorbed to dissolve the whole of the calcium chloride.

Other substances that are used for drying gases, *e.g.* phosphorus pentoxide or sulphuric acid, are also found to increase in weight when left exposed to the air.

One method for determining the amount of water in the air is to pass a known volume of the air through a weighed tube containing calcium chloride. The increase in weight of the calcium chloride tube indicates the amount of water in the volume of air used. If the air, after being dried, is passed through a weighed tube containing an alkali, *e.g.* potassium hydrate, or lime, the increase in weight in the tube will give the amount of carbon dioxide in the given volume of air.

Small amounts of other substances, such as ammonia, nitric acid, and compounds of sulphur, are found in the air, and within the last few years a remarkably inert substance to which the name of "argon" has been given, has also been discovered to be present.

Diffusion of Gases.

We shall first repeat some experiments we made when studying hydrogen and carbon dioxide.

To show that hydrogen can make its way into the air out of a tube held mouth downwards.

EXPT. 146. Fill several test-tubes with hydrogen; leave them in the stand, mouth downwards, but arranged so that the mouth

of the test-tubes are not entirely closed. Then, after allowing them to remain for varying periods—the first for one minute, the second for two minutes, and so on—test them in succession by bringing them, still mouth downwards, to a light. It will be found that tubes left for one or two minutes show, by the explosion that is produced on bringing them to a light, that some hydrogen has escaped, and air has taken its place. Tubes left for five minutes or longer give no result, showing that all the hydrogen has escaped.

EXPT. 147. **To show that carbon dioxide can make its way into the air out of a tube held mouth upwards.**

Fill a few tubes with carbon dioxide, and leave them for varying periods, mouth upwards. Then test at intervals of five minutes with lime-water. It will be found that for twenty minutes or thereabouts carbon dioxide can be detected, although in diminishing amounts. In a tube that has been standing about 25 minutes no milkiness is produced with lime-water, showing that all the carbon dioxide has escaped.

These results are quite different from what we might have expected. In the study of hydrogen, many experiments were described which proved hydrogen to be lighter than air. Among others we poured hydrogen *upwards* from one vessel to another. Yet we now find that hydrogen has actually made its way *downwards*, and out of the tube, air having at the same time taken its place. But the action in this case is only a slow one. In pouring the gas upwards the action is over in a second or two, whereas considerable time is required for the escape of the gas downwards. Carbon dioxide also was proved to be heavier than air, and yet we find now that it can make its way upwards out of a tube, and air make its way downwards, although here again considerable time is required.

This phenomenon of the mixing of gases when in contact, even though the heavier gas is below and the lighter one above, is called **Diffusion.** It is a phenomenon in which the separate particles, or *molecules*, of the gases are concerned. Each molecule is in a state of movement, and so it comes about that if two gases are in contact, the molecules of each wander off among the

molecules of the other, until a uniform mixture is the result. In our experiment the whole of one gas was contained in a test-tube, while the other was the air of a room, so that the amount of the first gas after mixing, in a small volume of the mixture, was quite inappreciable. Hence in the test-tubes we failed to notice any hydrogen or carbon dioxide at all.

To show the diffusion of oxygen and hydrogen through a glass tube.

The diffusion of oxygen and hydrogen may be shown in the following way :

Two soda-water bottles are filled, one with oxygen and the other with hydrogen, and connected by a piece of glass tubing passing through well-fitting corks. The apparatus is placed in a vertical position, the bottle filled with hydrogen being above. After an hour or two, it will be found that on applying a light to the mouth of each bottle, an explosion results, showing that the heavier oxygen has made its way upwards, and the lighter hydrogen has made its way downwards.

Fig. 65.

To show the diffusion of bromine vapour into the air.

The slow rise of a heavy gas in air may be *seen* in the case of bromine vapour. Bromine (see p. 207) is a heavy red liquid, which at the ordinary temperature rapidly changes into a red gas. A small quantity of bromine is passed down a thistle funnel, to the bottom of a tall cylinder (Fig. 65). The vapour is generated at the bottom, whence it can be seen to slowly make its way upwards, and in a very short time it will have risen to the top.

Although the air is composed of various gases of different

densities, we do not find it separating into layers in accordance with those densities, and this for two reasons. In the first place we have the phenomenon of **diffusion**, and in the second place we have the transfer of larger masses of air from one locality to another by **winds**. Both these causes also tend to prevent the accumulation, in excessive amounts, of carbon dioxide and noxious gases in large towns.

Gases will also diffuse through the fine holes in a piece of unglazed earthenware, or a plate of dried plaster of Paris. The following experiment shows this, and it also shows that different gases diffuse at different rates.

To show that hydrogen and air diffuse through a porous plug at different rates.

A glass tube *a*, 12 or 15 inches long, is closed at one end with a plug of plaster of Paris (Fig. 66). It is then filled with hydrogen or coal-gas by upward displacement, and the open end is put at once below the surface of water, in the beaker *b*. The rise of the water in the tube is sufficiently rapid to be apparent to the eye. The light gas is evidently diffusing out through the porous plug more rapidly than the air can diffuse in.

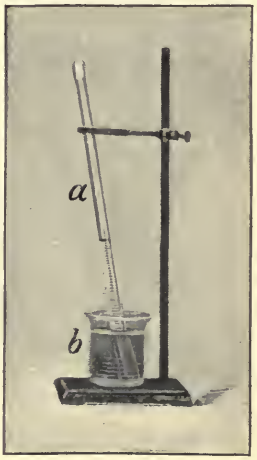

Fig. 66.

Graham, who carefully investigated the phenomenon of diffusion, formulated the following simple law: **Gases diffuse at rates which are inversely as the square roots of their densities.**

Thus oxygen, which is 16 times as heavy as hydrogen, diffuses at only one-fourth the rate. A given volume of oxygen takes four times as long as the same volume of hydrogen to diffuse through the same porous plug. It will have been noticed in Expts. 146 and 147, that it requires a much longer time for all

the carbon dioxide to diffuse out of a tube, than is required by hydrogen to diffuse out of a similar tube.

The different rates of diffusion of oxygen and hydrogen can be shown by passing an explosive mixture of these two gases through a clay tobacco-pipe. So much more rapidly does the hydrogen diffuse out through the porous walls, that the gas which may be collected, after passing through the pipe, will rekindle a glowing splint.

QUESTIONS. XIX.

155. Write a summary of the properties of nitrogen.

156. Describe two methods of obtaining nitrogen from air.

157. How is it proved that the oxygen and nitrogen in air are merely mixed, and not chemically combined?

158. How would you show that when a substance burns in air, it is only the oxygen with which it combines?

159. Give an account of the constituents of the atmosphere, distinguishing between its chief components, and those which occur in small quantities.

160. State the chief effects of animal and vegetable life on the atmosphere.

161. In what respects does the air in a crowded room differ from fresh air?

162. What is the composition of atmospheric air by volume? Air extracted from water is different in composition from ordinary air; to what extent, and why is it different?

163. State the law of diffusion of gases. If 10 c.cs. of hydrogen diffuse through a porous plate in 4 minutes, how long will 25 c.cs. of carbon dioxide (density 22) require to diffuse through the same porous plate?

164. A particular gas has a diffusion rate of 0.81, referred to air = 1.0. Find the density of the gas referred to hydrogen = 1.

165. How is the composition of air by weight ascertained?

166. Purified atmospheric air is passed slowly over metallic copper wire heated to redness. Supposing the copper to acquire 3.55 grams in weight, what weight of nitrogen would be left unabsorbed?

CHAPTER XX.

NITRIC ACID.

WHEN Cavendish exploded mixtures of oxygen and hydrogen, in his investigation of the composition of water, he found that the water formed gave, in some instances, an acid reaction. This happened when the oxygen used was not pure, but contained a small quantity of nitrogen, which was caused under the influence of the electric spark, to combine with oxygen. The product, when dissolved in water, gave nitric acid.

The union of oxygen and nitrogen can be shown by passing sparks from an induction coil between two platinum wires, about $\frac{1}{2}$ in. apart, in a dry flask full of air. The gas is seen to assume a yellow or brown colour. On introducing a little water and shaking, this colour disappears, and on testing with litmus the solution is seen to be acid.

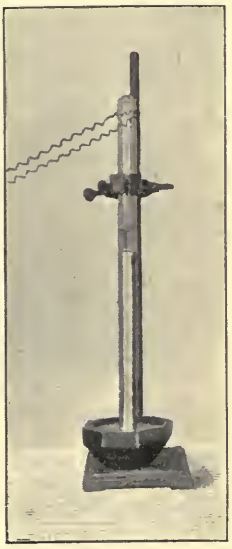

Fig. 67.

By a slight modification of this experiment, the whole of the nitrogen in a given volume of air may be made to unite with oxygen. A eudiometer is filled with mercury, and a certain amount of air is introduced (Fig. 67). A small quantity of a strong solution of potassium hydrate is passed up the tube, so as to float on the surface of the mercury. When sparks are passed between the platinum terminals, the brown gas that is formed unites with the alkali, and the volume is seen to diminish. There is not sufficient oxygen in air to combine with

the whole of the nitrogen, and so, when no further diminution in volume is seen to occur, more oxygen is introduced and the sparking is continued. In this way the whole of the nitrogen is removed, and the residual oxygen can be absorbed by passing a little strong solution of pyrogallol up the tube. Cavendish more than 100 years ago noticed that in an experiment similar to this there was always a bubble of gas—less than 1 per cent. of the original volume—left. In 1894 Lord Rayleigh showed that this was a hitherto unknown constituent of the air, which he called "argon."

Nitric acid is easily made from *nitrates*, two of which, potassium nitrate, called also nitre or saltpetre, and sodium nitrate, commonly called Chili saltpetre, are found in nature.

EXPT. 148. **To prepare nitric acid from potassium nitrate.**

Introduce about 30 grms. of potassium nitrate into a retort, and arrange it for distillation as in Expt. 23. Pour an equal weight of strong sulphuric acid down a funnel, into the retort, and then distil, using a small flame.[a] Notice the changes in appearance of the gas given off during the distillation, which must be stopped when 10 or 12 c.cs. of liquid have collected. The liquid is seen to have a yellow colour. This is owing to some of the brown fumes having dissolved in the acid which is itself colourless.

Pour the liquid left in the retort, while it is still hot, into a dish. On cooling it becomes solid.

EXPT. 149. Pour two or three drops of the distillate—not more—into a test tube. Fill up with water, shake well; take a drop of the solution on the tip of the finger and taste it. It has a sour or *acid* taste.

EXPT. 150. Add a few drops of blue litmus to the dilute solution. It is turned red.

In Expt. 38 it was found that nitric acid attacked many metals, but in no case was hydrogen given off. In most cases the metals are dissolved, nitrates being produced, and these are all soluble.

EXPT. 151. It would be well for the student to try again the effect of nitric acid on common metals, *e.g.* copper, lead, tin, iron, aluminium.

On account of its great activity, nitric acid received the name of *aqua fortis*, a name which is still sometimes used. Gold and platinum are exceptional in not being attacked by nitric acid; tin is attacked, but yields a white insoluble powder; while iron is exceptional in not being affected by the strong acid, although the dilute acid attacks it vigorously. Aluminium also is practically unaffected by nitric acid. A mixture of nitric and hydrochloric acids, called *aqua regia*, dissolves gold readily (see p. 205).

If nitric acid is heated strongly it decomposes, and among other things, free oxygen is produced.

EXPT. 152. **To show the production of oxygen by heating nitric acid.**

Support a churchwarden tobacco-pipe in a slanting position, with the end of the stem passing under water in a pneumatic

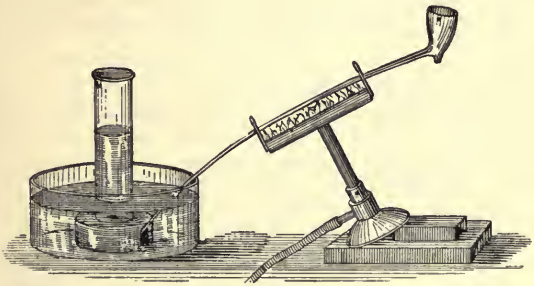

Fig. 68.

trough. When part of the stem has been made red-hot, pour nitric acid drop by drop into the bowl of the pipe. A colourless gas can be collected over the water in the trough, and on testing with a glowing splint, will prove to be oxygen.

On account of the ease with which nitric acid gives up its oxygen, it acts as a good oxidizing agent.

EXPT. 153. Heat a little sawdust on an iron tray, until it is

beginning to char. Then allow a drop or two of nitric acid to fall on it. It takes fire at once.

Unless great care has been exercised, a little acid will probably have touched the skin in the previous experiments. This will be indicated by a yellow stain. Nitric acid also attacks colouring matters; if the acid is allowed to touch the clothes, the colour is first destroyed, and if left long enough, the acid will destroy the material.

EXPT. 154. Add a little nitric acid to a solution of indigo. The colour is at once destroyed, a straw-coloured solution being obtained.

Nitrates.

The general method of making nitrates is to neutralize nitric acid with the proper base, and some have already been prepared. Thus lead nitrate is obtained by neutralizing nitric acid with litharge.[b] Nitric acid has also been found to attack most metals, and in nearly all cases a nitrate is the result. The blue liquid obtained by dissolving copper in nitric acid, on being evaporated, yields blue crystals of copper nitrate (see p. 148).

All nitrates are soluble in water.

They all yield nitric acid when heated with strong sulphuric acid.

When heated alone nitrates decompose, and most of them leave an oxide of the metal.

EXPT. 155. Heat a little calcium nitrate. Water is given off at first, and after a little time, brown fumes are evolved. When the residue is cold, add a little water. Notice the evolution of heat when the first few drops are added. Then add more water and test the solution by adding red litmus. It is turned blue. Blow through the solution for a short time. A white precipitate is produced. Clearly we have made calcium oxide by heating the nitrate.

EXPT. 156. Heat a little copper nitrate in a dry tube. It gives off water of crystallization, and, on further heating, decomposes, brown fumes being evolved, and black oxide of copper left.

EXPT. 157. Heat a little lead nitrate in a dry tube. Brown fumes are evolved, and yellow lead oxide, or litharge, is left.

Tests for nitrates.

To detect a nitrate, we first make from it nitric acid, and allow this to act on copper.

EXPT. 158. Add strong sulphuric acid to a little sodium nitrate, and then a few copper turnings. The nitrate and the sulphuric acid react and give nitric acid, which acts on the copper, and brown fumes are produced.

Another test for nitrates is made in the following way :

EXPT. 159. Add a cold solution of a nitrate to a cold solution of ferrous sulphate. Then pour strong sulphuric acid cautiously down the inside of the tube till it forms a separate layer at the bottom. A brown layer forms at the junction of the acid and the solution of the two salts. This is the "brown ring test" for nitrates.

Nitrates, like nitric acid, contain a good deal of oxygen, which is easily given up. They are therefore good oxidizing agents.

EXPT. 160. Heat a piece of charcoal with a blow-pipe flame until it is red-hot. Then drop a little potassium nitrate on it. The charcoal rapidly burns, or "deflagrates," as it is usually termed. This is also used as a test for nitrates.

EXPT. 161. Heat a little potassium nitrate in a hard glass tube until it melts and begins to effervesce. Then drop a small piece of sulphur into it. The sulphur burns away rapidly, obtaining its oxygen from the potassium nitrate, and giving a dazzling light.

It is on account of its oxidizing power that potassium nitrate is used in making gunpowder. Gunpowder is an intimate mixture of potassium nitrate, sulphur, and charcoal, and when once the two latter begin to burn, they are able to obtain all the oxygen they require from the potassium nitrate.

EQUATIONS.

(*a*) $KNO_3 + H_2SO_4 = KHSO_4 + HNO_3$.

(*b*) $PbO + 2HNO_3 = Pb(NO_3)_2 + H_2O$.

QUESTIONS. XX.

167. Write a summary of the properties of nitric acid.

168. Describe with a sketch the laboratory method of preparing nitric acid.

169. What effect is produced by pouring nitric acid on the following metals: copper, tin, lead, iron, aluminium?

170. Give examples of the oxidizing action of nitric acid.

171. Describe exactly how you would prepare a crystallized specimen of lead nitrate.

172. A white crystalline solid is supposed to be a nitrate. How would you proceed to find whether this is so or not?

173. What is the function of nitre in gun-powder? How would you obtain a specimen of nitre from gun-powder?

174. How would you demonstrate the presence of nitrogen in nitric acid?

175. How would you demonstrate the presence of oxygen in nitric acid?

176. Under what circumstances may oxygen and nitrogen be made to combine?

CHAPTER XXI.

AMMONIA. OXIDES OF NITROGEN.

EXPT. 162. Heat a little of the white crystalline solid ammonium chloride, or, as it is often called, sal-ammoniac, in a dry test-tube. It does not melt, but a white solid is seen to collect on the upper part of the tube, while the substance at the bottom gets less and less. At last nothing is left at the bottom of the tube but a small quantity of impurity. We say the ammonium chloride has **sublimed**, the substance formed on the side of the tube is a **sublimate**, and the whole operation is known as **sublimation.**

EXPT. 163. Notice that sal-ammoniac has little or no odour.

Then heat a little with caustic soda solution.[a] A pungent-smelling gas is given off. Hold a piece of moistened red litmus paper near the mouth of the tube. It is at once changed to blue. Hold a glass rod, moistened with hydrochloric acid, near the mouth of the tube. A dense white cloud is formed.

EXPT. 164. Heat a little sal-ammoniac with lime,[b] and repeat the above experiments. Similar results are obtained. The gas obtained in these experiments is called **ammonia.** The results obtained in the various experiments constitute *tests* for ammonia.

EXPT. 165. **To collect a little ammonia,** arrange the apparatus shown in Fig. 69. Make a mixture of sal-ammoniac and twice its bulk of lime, and place it in a test-tube, which is then to be fixed in a clamp with its mouth turned *slightly* downwards. Heat the mixture gently by moving the bunsen flame to and fro along the tube. The gas is to be collected in dry test-tubes by the method of upward displacement.

Fig. 69.

EXPT. 166. **To show that ammonia is lighter than air.**

Try to pour ammonia upwards from one tube to another. Its presence in the upper tube can readily be detected by the sense of smell.

EXPT. 167. Introduce a lighted splint into a tubeful of the gas. The splint is extinguished, but the gas itself appears inclined to burn, although nothing beyond a flickering brown

flame, and that only for an instant, is seen. In oxygen, however, ammonia burns readily with a brownish-yellow flame.

To show that ammonia will burn in oxygen.

The apparatus shown in Fig. 70 is convenient for showing this. The wide glass tube *a* is closed at the bottom with a cork, through which two tubes, each bent once at a right angle, pass.

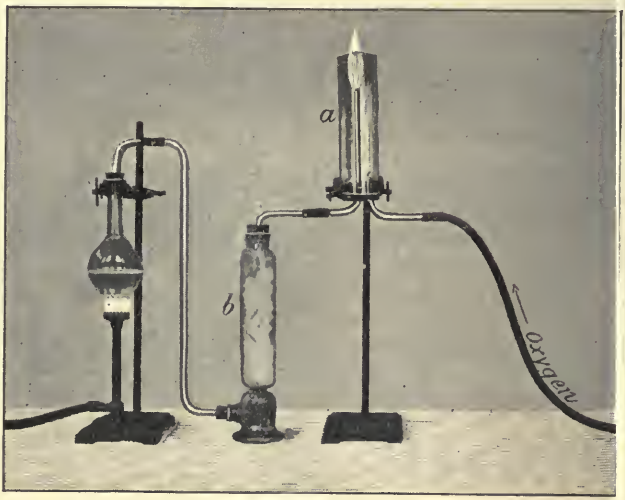

Fig. 70.

Ammonia passes through the longer one, which reaches nearly to the top of the wide tube. The ammonia is generated by warming a strong solution of the gas in water, and dried by passing through the glass tower *b* containing quick-lime. Through the short tube oxygen from a gas cylinder passes, and is distributed by passing through a layer of cotton-wool resting on the cork. On applying a light, the ammonia burns at the end of the long tube, with a brownish flame. On holding a cold bell-jar over the flame, moisture is deposited. Hence we see that hydrogen is one of the constituents of ammonia.

EXPT. 168. **To show the solubility of ammonia in water.**

Place another tube filled with the gas, mouth downwards, under water coloured with litmus, just reddened with acid. Shake the tube gently, without allowing any of the ammonia to escape. The water rises rapidly, and if all the air has been expelled it will fill the tube; the colour of the litmus will at the same time be changed to blue.

Ammonia is exceedingly soluble in water, one volume of water at the ordinary temperature dissolving about 800 volumes of the gas. The strongest solution has a specific gravity of 0.880. Ammonia gas is given off very readily on warming this solution. That ammonia is very soluble in water is shown in a very striking way in the following experiment.

The round-bottomed flask (Fig. 71) is filled with ammonia, and closed with a rubber stopper. Through a hole in this a glass tube passes, which tube can be closed either by a tap or by a piece of rubber tubing and a clip. The flask is placed so that the end of the glass tube passes nearly to the bottom of a vessel containing water coloured with red litmus. The tube is then opened and the flask cooled by pouring a little ether over it. The water rises slowly through the narrow tube, but when it reaches the top, and comes into contact with the bulk of the gas, the latter is almost instantly dissolved. A vacuum is thus produced, and then the water is violently forced up the tube by the pressure of the atmosphere, giving the appearance of a fountain in the flask.

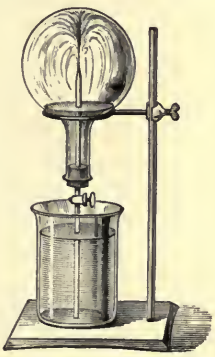

Fig. 71.

In its chemical properties the solution of ammonia in water resembles the solution made by dissolving sodium oxide in water, or sodium hydrate (see p. 54). It is hence often called ammonium hydrate.[c] On account of its alkaline properties we cannot use sulphuric acid for drying the gas. Ammonia unites with

calcium chloride, and hence this cannot be used as a drying-agent. Quick-lime is generally employed.

To prepare ammonium chloride.

If two jars, one containing ammonia gas and the other containing hydrochloric acid gas,[a] be brought together mouth to mouth, a white cloud of ammonium chloride is made. (Fig. 72.) The same substance may also be prepared in the following way.

EXPT. 169. Pour half a test-tubeful of dilute ammonium hydrate solution into a porcelain dish, and then add from a test-tube dilute hydrochloric acid, a little at a time. Stir continuously, and after each addition of acid, test a drop of the solution on litmus paper. Presently it will be found that the solution is without action on litmus. At this stage the pungent odour of ammonia is lost, and if a little of the solution be poured into a test-tube containing a little zinc no effervescence is produced. Now boil the solution to dryness. A white solid is left.

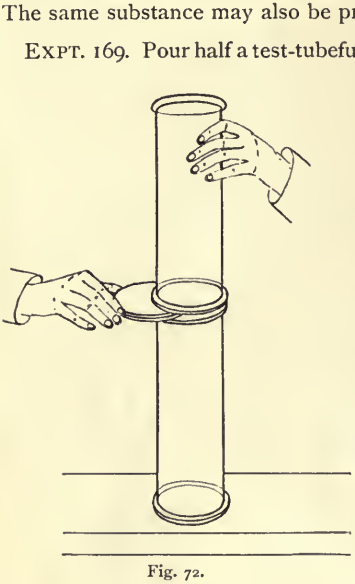

Fig. 72.

EXPT. 170. Heat a little of the solid obtained in the last experiment in a dry tube. A white sublimate is obtained.

EXPT. 171. Heat another portion with sodium hydrate solution. The odour of ammonia is at once noticed.

The white solid obtained is ammonium chloride, or sal-ammoniac. Other salts of ammonia, *e.g.* ammonium sulphate and ammonium nitrate, may be made in the same way, using the proper acid.

When coal-gas from the retorts is cooled, one of the substances obtained is an aqueous liquid, containing ammonia and other substances. This is the "ammonia liquor of the gas-works," and is the chief source of ammonium compounds. The liquor is heated with lime, and the ammonia which is given off is passed into water to form ammonium hydrate, or into acids to form ammonium salts. Ammonium chloride and ammonium sulphate are the salts chiefly prepared.

The **composition of ammonia** gas is shown by the following experiments :

1. **To show that the constituent gases occupy twice the volume of the ammonia.**—Dry ammonia is put in the eudiometer, and the volume under atmospheric pressure is

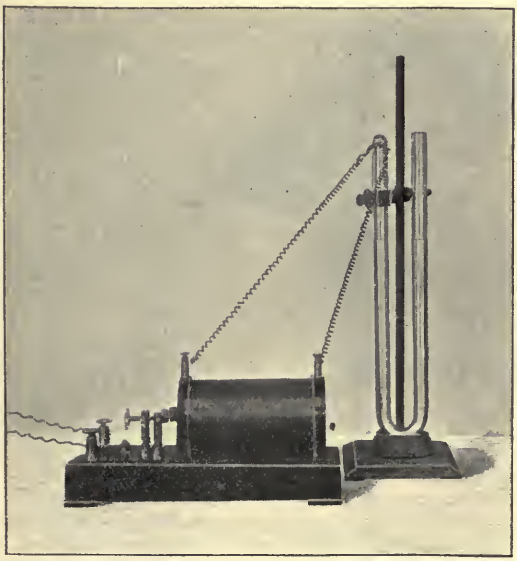

Fig. 73.

measured. On passing sparks through the ammonia it is decomposed, and it is found that the volume increases until

it is doubled. No greater increase than this can be obtained, and so we see that by decomposing ammonia the constituents occupy twice the volume of the original gas.

2. **To show that in ammonia one volume of nitrogen is combined with three volumes of hydrogen.**—A long tube divided into three equal parts by rubber bands, is closed at one end, and connected at the other with a tap funnel, which is either blown on to the tube or passes through a well-fitting cork. The tube is filled with **chlorine** gas, and then a strong solution of ammonia is admitted from the tap funnel, a drop at a time. As the first few drops enter, a greenish flame is produced, and ultimately the whole of the chlorine will disappear.

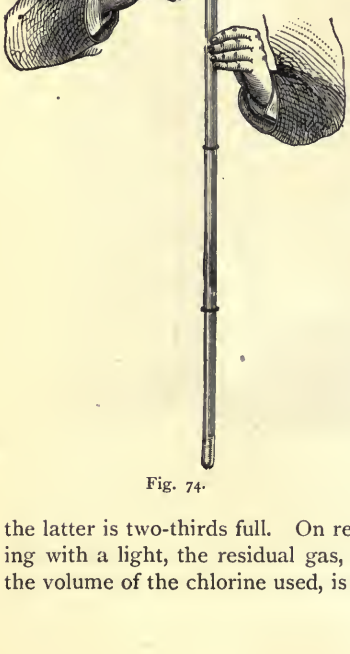

Fig. 74.

The chlorine unites with the hydrogen which we know to be present in ammonia, and, as will be seen later, a given volume of chlorine always unites with its own volume of hydrogen. Hence, whatever the amount of ammonia that has been acted on by the chlorine, it must have contained a volume of hydrogen sufficient to fill the tube. The funnel is now filled with dilute sulphuric acid, and one end of a bent tube is fitted into the mouth of the funnel, the other end passing below the surface of dilute sulphuric acid contained in a beaker. On opening the tap the sulphuric acid is drawn into the tube, until the latter is two-thirds full. On removing the stopper, and testing with a light, the residual gas, in volume equal to one-third the volume of the chlorine used, is found to be nitrogen.

Hence, we see that in ammonia three volumes of hydrogen are combined with one volume of nitrogen. Now we have already found that the constituent gases occupy twice the space of the original ammonia, and therefore *two volumes of ammonia contain one volume of nitrogen, combined with three volumes of hydrogen.*

Oxides of Nitrogen.

There are no fewer than five compounds of nitrogen and oxygen, but of these we shall study only three.

Nitrous oxide. For the preparation of this we require the salt, ammonium nitrate.

EXPT. 172. Heat a little ammonia solution in a porcelain dish. Both ammonia gas and water are given off, and there is practically no residue.

Heat a little dilute nitric acid in the same way. Here again, on evaporating to dryness, we have no residue left. The production of brown fumes shows that the nitric acid is, to some extent at any rate, being decomposed as it is heated.

Now proceed to neutralize nitric acid with ammonia solution.[e] Put about half a test-tubeful of dilute ammonia solution (1 part strong solution to 3 parts water) in a porcelain dish. Add dilute nitric acid from another tube, a little at a time, stirring after each addition, and testing a drop on the end of the glass rod, with litmus paper. It will be found ultimately that a neutral liquid, without odour, has been produced. Boil this to dryness. A white solid is left. This is **ammonium nitrate.**

EXPT. 173. Heat a little ammonium nitrate in a small tube, as strongly as possible. Hold the mouth of the tube slightly downwards. The salt melts, and then effervesces violently. Much steam will be given off, and this renders it impossible to test the gas which accompanies it. Towards the end it will be found that the salt decomposes with a slight explosion.

EXPT. 174. To collect some of the gas given off when ammonium nitrate is heated, we require an apparatus similar to

that shown in Fig. 75. Put about 20 grm. of the salt in the flask, and heat with a small flame. *Do not hold the bunsen burner in the hand.* The gas is rather soluble in cold water, and so it is often collected over warm water. Cold water may be used, but there is considerable loss in this case. After allowing the gas to escape for a little while, fill one or two small test-

Fig. 75.

tubes with the gas, and then proceed to fill three jars. It will be seen that the gas is colourless, and is not altered when brought into contact with air.

EXPT. 175. Put a little water into a tubeful of the gas, cover tightly with the thumb, and shake vigorously. Then open the tube under water, when it will be found that the water rises, showing that some of the gas has been dissolved. Shake again, open under water again, and repeat a number of times. By this means the greater part of the gas may be dissolved. One volume of water at 0° C. dissolves 1·3 volumes of the gas.

EXPT. 176. Introduce a glowing splint into the gas. It is rekindled and burns vigorously. The gas might be oxygen.

EXPT. 177. Introduce a piece of feebly burning sulphur into a jar of nitrous oxide. The sulphur is extinguished. This result is different from the corresponding result with oxygen. Cause the sulphur to burn brightly and again introduce it into the gas. It continues to burn still more brightly.[g] Cautiously test the odour of the gas left in the jar. It is that of "burnt sulphur' Shake with water and add blue litmus. The colour changes to red.

EXPT. 178. Introduce a piece of burning phosphorus. It continues to burn, but still more brightly, and a white solid is formed. This is dissolved by water, and the solution reddens blue litmus.

We can show that nitrous oxide contains nitrogen as well as oxygen, by passing it through a tube containing heated copper.[h] Oxide of copper is seen to be produced, and the gas that passes on may be collected over water. On testing it proves to be nitrogen. Still another distinction from oxygen will be noticed later (see p. 149).

Nitrous oxide is commonly called "laughing gas." It produces temporary insensibility when breathed, and is the gas used by dentists for this purpose.

Nitric oxide. We shall next try to collect some of the reddish-brown gas formed when copper is dissolved in nitric acid. An apparatus similar to that used in collecting hydrogen will be used.

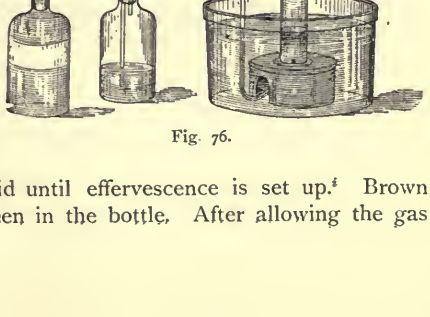

Fig. 76.

EXPT. 179. Place about 20 grm. of copper turnings in the bottle; add a little water, and then strong nitric acid until effervescence is set up.[i] Brown fumes are at once seen in the bottle. After allowing the gas

to escape for a little while, fill two or three jars by displacement over water. The gas is seen to be colourless. It is almost insoluble in water.

By boiling the solution left in the bottle to a small bulk, and allowing to cool, blue crystals of copper nitrate are obtained.

EXPT. 180. Uncover one of the jars so as to allow air to come in contact with the gas. Brown fumes are at once produced.[1] Add water and shake. The brown fumes dissolve. Add a little blue litmus. It is turned red.

EXPT. 181. Introduce a lighted splint into the second jar. It is extinguished, and the gas itself does not burn.

EXPT. 182. Introduce a piece of phosphorus, which is burning only feebly, into a jar of the gas. It is extinguished. Then take out the phosphorus and introduce it again when burning briskly. It continues to burn very brightly, the appearance being similar to that observed when phosphorus burns in oxygen or in nitrous oxide. Add a little water and shake. The white fumes dissolve. Test the liquid with blue litmus. It is reddened. It appears therefore as though the phosphorus

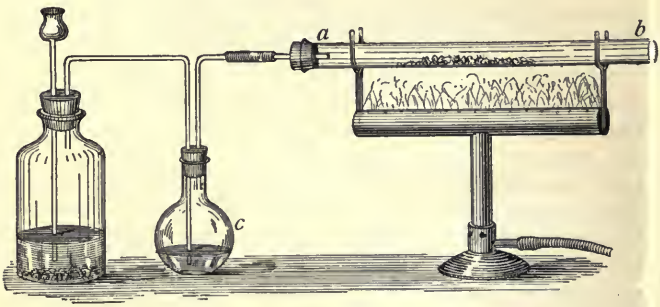

Fig. 77.

united with oxygen when burnt in nitric oxide. We may try in another way whether oxygen is one of the constituents of nitric oxide.

A glass tube containing copper turnings is heated, and nitric oxide is passed through it, exactly as in the case of nitrous

oxide. The gas which passes out of the tube may be collected over water. On testing, it proves to be nitrogen. The copper is seen to be coated with copper oxide, and so we see that nitric oxide consists of nitrogen and oxygen.

We have learnt, therefore, that nitrous oxide, nitric oxide, and air all consist of nitrogen and oxygen, but it can be shown that, whereas the proportions of oxygen and nitrogen in air vary slightly in different places, the proportions of these elements in nitric oxide and in nitrous oxide are invariable. We conclude, therefore, for this and other reasons, that whereas in air the oxygen and nitrogen are merely mixed, in nitrous oxide and nitric oxide they are chemically combined.

The action of oxygen on nitric oxide.

EXPT. 183. Half fill a test-tube with air, then pass nitric oxide into it, from another tube, a little at a time. It will be found that the brown fumes which are formed dissolve in the water, and presently the addition of another bubble of nitric oxide forms no more red fumes. If the gas that is left be tested, it will be found to be nitrogen. The water left in the tube will turn blue litmus red.

EXPT. 184. Half fill a test-tube with oxygen, and pass nitric oxide in as before. The brown fumes which are momentarily visible, dissolve in the water, and the volume of gas gets less and less until only a very small amount remains. If both gases were pure, no gas at all would be left.

The brown gas which dissolves in water giving an acid liquid is called **nitrogen peroxide.** From the mode of formation, it follows that it contains a greater proportion of oxygen than nitric oxide. It is produced when many nitrates, *e.g.* lead nitrate, are heated alone. It is the oxygen only in the air which forms brown fumes with nitric oxide. Nitrous oxide has been shown to contain oxygen and to resemble oxygen in many respects. If, however, nitric oxide be passed into a tube containing nitrous oxide, no brown fumes are formed. Nitric oxide is a test for uncombined oxygen only; if the oxygen is in a state of combination, *e.g.* in nitrous oxide, nitric oxide cannot combine with it.

The amount of oxygen in the air was formerly determined by absorbing it with nitric oxide; more convenient and accurate methods are now employed.

EQUATIONS.

(*a*) $NH_4Cl + NaOH = NaCl + NH_3 + H_2O$.
(*b*) $2NH_4Cl + CaO = CaCl_2 + 2NH_3 + 2H_2O$.
(*c*) $NH_3 + H_2O = NH_4OH$.
(*d*) $NH_3 + HCl = NH_4Cl$.
(*e*) $NH_3 + HNO_3 = NH_4NO_3$.
(*f*) $NH_4NO_3 = N_2O + 2H_2O$.
(*g*) $S + 2N_2O = SO_2 + 2N_2$.
(*h*) $Cu + N_2O = CuO + N_2$.
(*i*) $3Cu + 8HNO_3 = 2Cu(NO_3)_2 + 4H_2O + 2NO$.
(*j*) $NO + O = NO_2$.

QUESTIONS. XXI.

177. Write a summary of the properties of ammonia.

178. Sketch the apparatus you used in the preparation of ammonia. What materials did you employ in the preparation?

179. A given liquid is supposed to contain an ammonium salt in solution. How would you ascertain whether this was the case or not?

180. How would you prepare nitrogen from ammonia?

181. How would you show (1) that ammonia is very soluble in water, (2) that ammonia is lighter than air?

182. How would you demonstrate the existence of nitrogen and of hydrogen in ammonia?

183. What effect would be observed on introducing the following into a tube of ammonia standing over mercury: (*a*) water, (*b*) a piece of freshly-burnt charcoal, (*c*) hydrochloric acid gas?

184. Ammonia gas is passed into (1) water, (2) dilute sulphuric acid. Describe what takes place in each case. How could the gas be again obtained from each of these solutions?

185. What is the effect of boiling a solution of sal-ammoniac with lime or potash? Name the gas evolved, and state from what other sources it can be obtained.

186. Write a summary of the properties of nitrous oxide.

187. A glowing splint, introduced into a jar of gas, bursts into flame; what gases are capable of producing this result, and how are they distinguished from one another?

188. How would you obtain nitrogen from (1) nitrous oxide; (2) nitric oxide?

189. How is nitrous oxide distinguished from nitric oxide?

190. How would you prove the presence of oxygen in nitrous and nitric oxides?

191. Write a summary of the properties of nitric oxide.

192. Describe the method of making nitric oxide gas. What substance is left in the generating bottle, and how would you obtain it in a crystalline form? What would be observed on heating one of these crystals?

193. What is the effect of heating (*a*) lead nitrate, (*b*) potassium nitrate; (*c*) ammonium nitrate?

CHAPTER XXII.

WEIGHT OF A LITRE OF AIR. LAW OF CHARLES.

DEFINITE amounts of gases, as of solids and liquids, have definite weights. The operation of finding the weight of any given volume of gas with accuracy is one of extreme difficulty. An approximate result in the case of air can be obtained by the following method:

EXPT. 185. **To find the weight of a litre of air at the temperature of the laboratory.**

A round-bottomed flask of 300-400 c.cs. capacity is provided with a well-fitting rubber stopper, through which passes a short piece of glass tube. This can be closed by a piece of rubber tubing and a clip (Fig. 78). Place 40-50 c.cs. of water in the flask, put the stopper in very securely, leave the tube open, and heat the water.

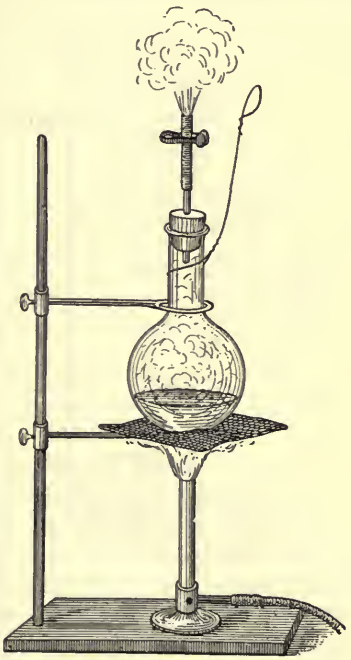

Fig. 78.

After it has been boiling briskly for about 10 mins., all the air will have been expelled, and the flask will be full of steam. Close the tube with the clip, and *at once* remove the flame. Wipe the outside of the flask, allow it to cool, and then weigh. Cautiously open the clip to allow air to enter, and weigh again. The increase represents the weight of the volume of air which enters. This volume is obtained by (1) measuring the volume of water left in the flask, and (2) measuring the volume of water required to fill the flask to the lower end of the stopper, and subtracting the former from the latter. We can then calculate the weight of any volume, say 1 litre, of air at the temperature of the laboratory.

The results of an experiment by this method are given below:

Second weighing = 86.385 grm. = wt. of flask + water remain. + air.
First " = 86.066 " = wt. of flask + water remaining.
Difference = 0.319 grm. = wt. of air entering.

Capacity of flask = 296 c.cs.
Volume of water remaining in flask = 30 c.cs.
Volume of air = 266 c.cs.

This weighs 0.319 grm.

$$\therefore \text{ 1 litre weighs } \frac{0.319 \times 1000}{266} = 1.2 \text{ grm.}$$

More accurate results show that

1 litre of air	weighs	1.29	grm.	at 0° C. and 760 mm. pressure.
1 " hydrogen	"	0.09	"	
1 " oxygen	"	1.43	"	
1 " nitrogen	"	1.25	"	
1 " carbon dioxide	"	1.98	"	

The Densities of Gases.

It is found that hydrogen is lighter than any other gas, and it is, by chemists, taken as the standard of density for gases. The relative density of hydrogen is called 1, and the number expressing the relative density of any gas, shows how many times any volume of that gas is heavier than the same volume of hydrogen.

If x = the density of air.

Then $x \times 0.09 = 1.29$,

$$x = \frac{1.29}{0.09} = 14.3.$$

Similarly we obtain

$$\text{density of oxygen} = \frac{1.43}{0.09} = 16.$$

$$\text{density of nitrogen} = \frac{1.25}{0.09} = 14.$$

$$\text{density of carbon dioxide} = \frac{1.98}{0.09} = 22.$$

Expansion of gases by heat. The Law of Charles. Gases, like all other substances, expand on being heated, but in their case the expansion for a given rise in temperature is greater than in the case of solids or liquids.

EXPT. 186. Fit a small flask with a rubber stopper, through which passes a straight glass tube. Fix this in a retort stand in an inverted position, so that the end of the glass tube dips below the surface of some coloured water. Even the warmth of the hand is sufficient to cause the air in the flask to expand and drive some out through the coloured liquid. On allowing the flask to cool again, the air contracts, and liquid is drawn up the narrow tube to take the place of the air that was expelled.

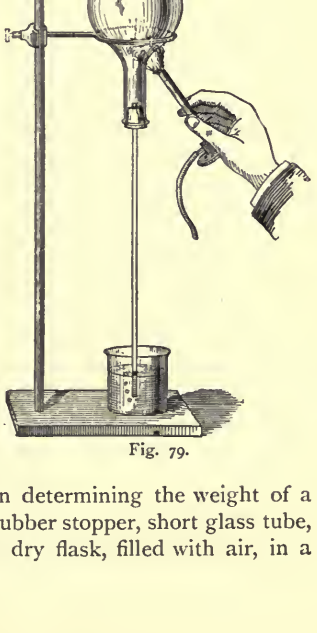

Fig. 79.

EXPT. 187. **To find the expansion of air on being heated from 0° C. to 100° C.** A flask similar to that used in determining the weight of a litre of air, and provided with a rubber stopper, short glass tube, and clip, is required. Place the dry flask, filled with air, in a

vessel of boiling water, the clip being open. Leave it for 10 minutes, by which time the air contained in the flask will have attained the temperature of the boiling water. Close the tube with the clip and remove the flask. Allow it to cool for a few minutes, then put the end of the tube under water, and cautiously open the clip. Water enters the flask, and after shaking for a little while, the flask and the air in it will be at the same temperature as the water. Take the temperature of the water, close the clip, remove the flask, and measure the volume of water which has entered.

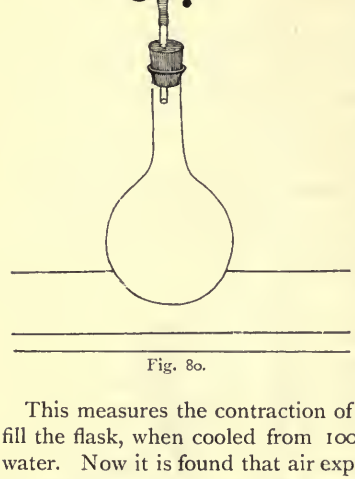

Fig. 80.

This measures the contraction of a volume of air, sufficient to fill the flask, when cooled from 100° to the temperature of the water. Now it is found that air expands or contracts very nearly uniformly as the temperature varies, and hence we may calculate what the contraction would be if the air were cooled down to 0°.

An experiment gave the following result:

Capacity of flask, - - -	296 c.cs.
Water entering on cooling, -	71 c.cs.
Temperature of cold water, -	12° C.

∴ Contraction between 100° and 12°, *i.e.* for 88° = 71 c.cs.

$$\therefore \text{ Contraction for } 100^\circ = \frac{71 \times 100}{88} = 80.7 \text{ c.cs.}$$

Hence 296 c.cs. at 100° would become 296 − 80.7 = 215.3 c.cs. at 0°

Or

215.3 c.cs. would expand by 80.7 c.cs. when heated from 0° to 100°.

∴ 1 c.c. „ $\frac{80.7}{215.3}$ c.cs. „ „

∴ 1 c.c. „ $\left(\frac{80.7}{215.3} \times \frac{1}{100}\right)$ c.cs. „ 0° to 1°.

i.e. 1 c.c. „ .00375 c.cs. between 0° and 1°.

The calculation would have been shorter if the cold water had contained melting ice, when the temperature would have been 0°.

The temperature 0° C. is often referred to as the "normal temperature."

The amount by which unit volume of a gas increases in volume when heated from 0° C. to 1° C. under constant pressure is called its **coefficient of expansion.**

If the flask be filled with hydrogen by upward displacement, and if, instead of keeping the tube open during the time the flask is being heated, it be momentarily opened at intervals during the operation, the coefficient of expansion of hydrogen may be determined. It is found to agree very closely with the coefficient of expansion of air.

It will be noticed that in the above calculation we assumed that the expansion was regular between 0° and 100°. This is approximately true for the majority of gases. It is found also, that gases have nearly the same coefficient of expansion. This is about .0366, or $\frac{1}{273}$. We may say then, that a given volume of any gas at 0° C., expands by $\frac{1}{273}$ of that volume for every degree rise of temperature.

Hence 1 c.c. at 0° becomes $(1+\frac{10}{273}$ c.c.) at 10°, $(1+\frac{100}{273}$ c.c.) at 100°, and $(1+\frac{273}{273})$, *i.e.* the volume would be doubled, at 273°. On the other hand, cooling below 0° causes a contraction. At −10° 1 c.c. of gas becomes $(1-\frac{10}{273})$, at −100°, $(1-\frac{100}{273})$, and $(1-\frac{273}{273})$, *i.e. nil*, at −273°. No gas varies in volume regularly down to this low temperature, in fact all known gases liquefy before this temperature is reached.

The temperature, −273° C., at which the volume occupied by any gas, supposing it to obey continuously the law of expansion of gases, would be nothing, is called the **absolute zero** of temperature. The boiling point of hydrogen is 35° abs. Any temperature on the Centigrade scale is transformed to the absolute scale by adding 273°. Thus water freezes at 273° abs., and boils at 373° abs.

The statement that

All gases expand by $\frac{1}{273}$ of their volume at 0° C.,

for every increase in temperature of 1°, the pressure being kept constant,

is generally referred to as the "Law of Charles." It was first announced at the beginning of the present century.

If we have unit volume of gas at 0° C., it will become

$$\left(1+\frac{t_1}{273}\right) \text{ at } t_1^\circ, \text{ and } \left(1+\frac{t_2}{273}\right) \text{ at } t_2^\circ,$$

if the pressure remains the same. This is true whether the sign of t_1 or t_2 is + or −. Hence

$$\frac{\text{Vol. at } t_1^\circ}{\text{Vol. at } t_2^\circ}=\frac{1+\frac{t_1}{273}}{1+\frac{t_2}{273}}=\frac{273+t_1}{273+t_2}.$$

But these expressions represent the absolute temperature of t_1°C and t_2°C respectively, hence we see that

The volume of a given quantity of gas is directly proportional to its absolute temperature if the pressure is constant.

This is another way of stating the Law of Charles.

We shall now work a few problems on the connection between volume and temperature of a given quantity of gas.

Example 1. *A certain quantity of gas occupies* 4 *litres at* 0° *C. By what amount will the volume increase when the temperature is raised to* 100° *C?*

The volume is directly proportional to the absolute temperature.

$$0^\circ \text{ C.}=273^\circ \text{ abs.},$$
$$100^\circ \text{ C.}=373^\circ \text{ abs.}$$

Suppose V=vol. at 0° and V_1=vol. at 100°.

Then $$\frac{V}{V_1}=\frac{273}{373},$$

i.e. $$\frac{4}{V_1}=\frac{273}{373}.$$

Hence $$V_1=\frac{4\times 373}{273}=5.5 \text{ litres.}$$

Hence the volume is increased by 1.5 litres.

Example 2. 30 *c.cs. of air at* 15° *C. are heated until the volume is doubled. To what is the temperature raised?*

Let $x°$ C. = the temperature C.

$15°$ C. = 288° abs.,

and $x°$ C. = $(x+273)°$ abs.

Suppose V = vol. at 15° C., and V_1 = vol. at $x°$ C.

Then $$\frac{V}{V_1}=\frac{288}{x+273}.$$

i.e. $$\frac{30}{60}=\frac{288}{x+273}.$$

$\therefore$ $$x+273=576.$$

Hence $x°$ C. = 303° C.

Example 3. *Two litres of gas at* 15° *C. are cooled to* −40° *C. Find the volume at the lower temperature.*

15° C. = 288° abs.

−40° C. = 233° abs.

Suppose V = vol. at 15° C., and V_1 = vol. at −40° C.

Then $$\frac{V}{V_1}=\frac{288}{233},$$

$$\frac{2}{V_1}=\frac{288}{233},$$

Hence $$V_1=\frac{466}{288}=1.6 \text{ litres.}$$

QUESTIONS. XXII.

194. Describe an experiment to show that air has weight.

195. How would you prove that carbon dioxide is, bulk for bulk, about 1½ times as heavy as air?

196. 20 c.cs. of hydrogen, measured at 12° C., are heated to 100° C., what increase in volume will be observed?

197. A litre of oxygen at 15° C. is (*a*) heated to 25° C., (*b*) cooled to −25° C. What will be the volume in each instance?

198. At what temperature will the density of oxygen be the same as that of hydrogen at 0° C.?

199. If you have 150 c.cs. of dry carbon dioxide at 12° C., what volume will it occupy at 0° C.?

200. If 1 litre of hydrogen at 273° C. and under a pressure of 760 mm. weighs .045 grm., how many grms. will 1 litre of each of the following weigh at the same temperature: carbon dioxide, nitrogen, oxygen?

CHAPTER XXIII.

PRESSURE OF GASES. BOYLE'S LAW. CORRECTION OF VOLUMES FOR TEMPERATURE AND PRESSURE.

Torricelli's experiment. A glass tube more than 30 in. long, and sealed at one end, is filled with mercury. The other end is covered with the thumb, and put below the surface of mercury contained in a trough. On removing the thumb, it will be found that the whole of the mercury does not run out of the tube. It falls to a certain point and then becomes stationary. The length of the column remaining, measured from the surface of the mercury in the trough, will be about 30 in., or 760 mm. On tilting the tube it will be found that the mercury approaches the closed end of the tube, but the vertical height of the column from the level of the surface of the mercury in the trough will always be 30 in.

When the closed end is on a level with the surface of the mercury in the first position, the tube will be completely filled with mercury. Hence the space above the mercury in other positions can contain no gas. It is called "the Torricellian vacuum."

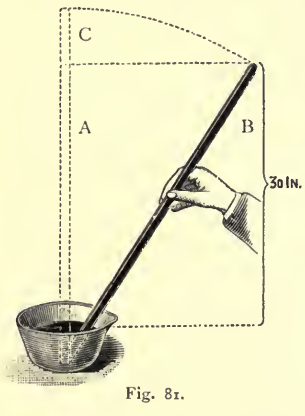

Fig. 81.

The mercury in the tube in the Torricelli experiment is kept in position by the pressure exerted by the atmosphere on the surface of the mercury in the trough. This pressure is not always the same. If such a tube be watched from day to day, it will be found that the height of the column varies, the variations corresponding with the variations in the pressure of the atmosphere.

An instrument which measures the pressure of the atmosphere is called a **barometer.**

To show that the height of the mercury column in the Torricelli experiment varies when the pressure on the mercury in the trough varies.

It may be shown by the following experiment that the height of the mercury column is increased when the pressure exerted by the air on the mercury in the trough is increased, and diminished when this pressure is diminished.

The "trough" in this case is a wide-mouthed bottle *a*, half full of mercury and provided with a tightly-fitting stopper. Through this the tube *b* containing mercury passes. The space above the mercury column is a Torricellian vacuum. The surface of the mercury in the bottle is in contact with the outer air through the tube *c*. If this be attached by means of a rubber tube to a filter-pump, the pressure on the surface of the mercury in the bottle is diminished, and the mercury falls in the tube. The same result may be obtained by sucking out some of the air through *c*. If, on the other hand, we blow through the tube *c*, the pressure is increased, and the mercury rises.

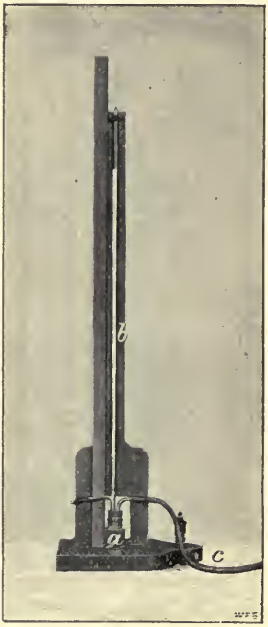

Fig. 82.

It has been mentioned already that the average height of the mercury column in a barometer is about 30 in., or 760 mm. A pressure equal to that exerted by a column of mercury 760 mm. high is called "a pressure of one atmosphere." This is usually called the "normal pressure."

If the tube in the Torricelli experiment have a sectional area

of 1 sq. in., then the pressure of the atmosphere on 1 sq. in. must be equal to the weight of 30 cubic in. of mercury. This is equal to a weight of 15 lbs., very nearly, and the "normal pressure," or the "pressure of one atmosphere," is therefore about 15 lbs. on a square inch.

To show that the volume of a gas diminishes as the pressure to which it is subjected is increased, the temperature remaining constant.

EXPT. 188. Take a tube of the shape shown in Fig. 83, the end of the shorter limb being closed. Introduce sufficient

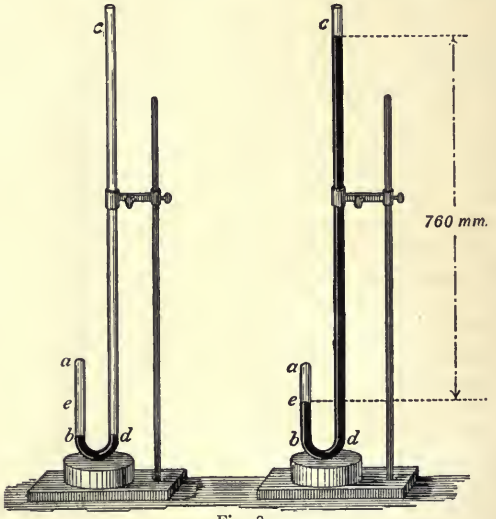

Fig. 83.

liquid, water or mercury, to form a column an inch or so long in each limb. By allowing a little air to pass out from the closed limb, it will be possible to bring the liquid to the same level in both limbs. Consider now the state of a thin layer of liquid at the lowest part of the tube. It will be at rest, and must therefore be subject to the same pressure on both sides.

On each side there is the pressure exerted by the mercury. As the height of the two columns is the same, the pressures exerted by them will balance one another. On the side towards the open limb there is in addition the pressure exerted by the atmosphere on the surface at *d*. On the other side, there must consequently be exerted by the air in the closed limb, a pressure equal to that of the atmosphere. If the surface *d* is raised by pouring more mercury into the open limb, it will be found that the surface *b* rises also, but less rapidly. The air offers a certain resistance to compression, but nevertheless it is seen that as the pressure on it is increased, the volume does diminish.

EXPT. 189. **To find the relation between the volume of a gas and the pressure to which it is subjected, the temperature remaining constant.**

Instead of altering the pressure by pouring mercury into the tube, the arrangement shown in Fig. 84 may be used. The glass tube *a*, closed at one end, and fixed to the stand, is connected at the open end by a piece of thick-walled india-rubber tubing, to the tube *b*, open at both ends. The tube *b* is attached to a piece of wood which slides in a groove, and which can be fixed in any desired position by a screw at the back of the apparatus. Mercury is poured in through the tube *b*, so as to imprison a certain volume of air, and the volume of air and the amount of mercury is so arranged that when the slider is in its extreme positions, *i.e.* at the top and bottom of the grooves, the surfaces of the mercury remain in the glass tubes. If we

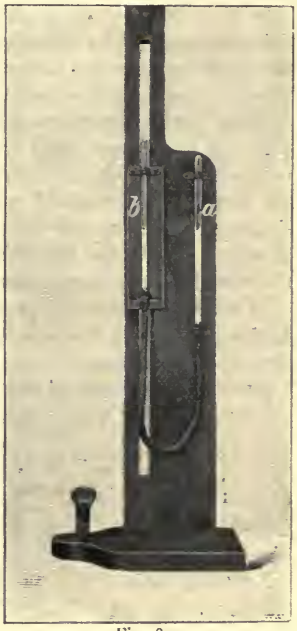

Fig. 84.

assume that the closed tube is of uniform bore throughout, the distance of the surface of the mercury in this tube from the closed end may be taken to represent the volume of air enclosed.

Arrange the movable tube so as to have the surfaces of the mercury at the same level. Measure the height of the closed end from the base of the apparatus, and also the height of the mercury in either tube. The difference between these readings may be taken to represent the volume of air in the tube *a* at atmospheric pressure. Read the barometer to find what that pressure is at the time of the experiment.

Now raise the movable glass tube, when it will be seen that the surface of the mercury in the tube *b* is above that in the tube *a*. Measure the height of both surfaces from the base of the apparatus. The difference between these readings represents the pressure over and above the pressure of the atmosphere, to which the confined volume of air is subjected. Continue to increase the pressure, reading the heights of the two surfaces of mercury each time until the movable tube is in its highest position. Next move it below its first position. The surface of the mercury in the tube *b* will now be below that in the tube *a*, so that the pressure on the air will be less than the atmospheric pressure. Make the same readings as before, and continue till the movable tube is in its lowest position.

It will be found in all cases, that the volume of air diminishes as the pressure on it increases, and as the pressure diminishes the volume increases.

Arrange the results in the following manner:

	Length of Air Column ($\propto V$).	Height of Mercury Columns.			Height of Barometer (B).	Total Pressure of Gas. $P=B+(L-R)$.	Product ($P\times V$).
		Left (L).	Right (R).	Difference ($L-R$).			
I							
etc.							

In the last column the value of the product (pressure × volume) is given

It is found that this product is the same for any given quantity of most gases.

$$P \times V = \text{a constant.}$$

If P were doubled V would be halved, if P were diminished to one-fourth V would become four times as great. We may state the result in another way, therefore, viz. that

The volume of any given quantity of gas varies inversely as the pressure, if the temperature remains constant.

This is generally known as "Boyle's Law," after its discoverer. It was first announced in 1662.

We can now easily calculate the volume of any given quantity of gas under different pressures.

Example 1. *A certain vessel contains* 5 *litres of air at normal pressure. What would be the volume of the same quantity of air under a pressure of* 700 *mm. of mercury, the temperature remaining constant?*

Here
$$P \times V = \text{a constant} = P_1 \times V_1;$$
$$\therefore \quad 760 \times 5 = 700 \times V_1;$$
$$\therefore \quad V_1 = \frac{760 \times 5}{700} = 5.43 \text{ litres.}$$

Example 2. 20 *c.cs. of air at normal pressure are compressed to* 13 *c.cs. What is the increase of pressure to which the air is subjected, the temperature remaining constant?*

Find first the total pressure, P_1, in the second case.

$$P \times V = P_1 \times V_1,$$
$$760 \times 20 = P_1 \times 13,$$
$$P_1 = \frac{760 \times 20}{13} = 1169 \text{ mm.}$$

But the original pressure was 760 mm., hence the increase

$$= 409 \text{ mm.}$$

Example 3. *A tube, partly filled with mercury, is inverted in a trough of mercury. The surface of the mercury in the tube is* 14 *cm. above the surface of that in the trough. What is the pressure, in atmospheres, exerted by the gas in the tube?*

The pressure exerted by the gas + the pressure exerted by a column of mercury 14 cm. or 140 mm. high, is equal to the atmospheric pressure, *i.e.* 760 mm. of mercury.

Hence the gas exerts a pressure equal to

$$(760 - 140) \text{ mm.} = 620 \text{ mm. of mercury.}$$

This $= \frac{620}{760} = \frac{31}{38}$ of an atmosphere.

To find the change in volume of any given quantity of gas when both the temperature and pressure vary.

If the pressure and temperature to which a certain quantity of gas is subjected be varied, we may proceed to find the volume under the new conditions, as in the following examples.

Example 1. *At* 12° *C. and under normal pressure, a certain quantity of a gas occupies* 400 *c.cs., what would be the volume at* 20° *C. and under a pressure of* 700 *mm. of mercury?*

First suppose the pressure remains constant at 760 mm., and correct for temperature.

$$12° \text{ C.} = 285 \text{ abs.},$$
$$20° \text{ C.} = 293 \text{ abs.}$$

Suppose V = vol. at 12°, and V_1 = vol. at 20°.

Then
$$\frac{V}{V_1} = \frac{285}{293},$$
$$\therefore \quad V_1 = \left(\frac{293 \times 400}{285}\right) \text{c.cs.}$$

= volume the gas would occupy at 20° and under a pressure o 760 mm.

Having obtained the volume the gas would occupy at th new temperature, but under the original pressure, proceed t correct this for the alteration in pressure.

$$P_1 V_1 = P_2 V_2,$$
$$760 \times \left(\frac{293 \times 400}{285}\right) = 700 \times V_2,$$
$$V_2 = \frac{760 \times 293 \times 400}{285 \times 700} = 446 \text{ c.cs.}$$

Example 2. *A certain quantity of gas occupies* 50 *c.cs. at* 15° *C. and* 730 *mm. pressure. What would be the volume at N.T.P.?*

(N.T.P. is a usual contraction for "normal temperature and pressure," *i.e.* 0° C. and 760 mm. pressure.)

Correct first for temperature,

$$15° \text{ C.} = 288° \text{ abs.},$$
$$0° \text{ C.} = 273° \text{ abs.}$$

Suppose V = vol. at 15°, and V_1 = vol. at 0°.

Then
$$\frac{V}{V_1} = \frac{288}{273},$$
$$V_1 = \left(\frac{273 \times 50}{288}\right) \text{ c.cs.}$$

Correcting this volume for pressure we get

$$P_1V_1 = P_2V_2,$$
$$730 \times \left(\frac{273 \times 50}{288}\right) = P_2V_2 = 760 \times V_2,$$
$$\therefore \quad V_2 = \frac{730 \times 273 \times 50}{760 \times 288},$$
$$\therefore \quad V_2 = 45.5 \text{ c.cs.}$$

QUESTIONS. XXIII.

201. How would you show that the volume of a gas varies (*a*) as the pressure, (*b*) as the temperature, varies?

202. State Boyle's Law. Calculate the pressure exerted by air confined in a tube over mercury when the mercury inside the tube stands 13 cm. above the mercury in the trough. (Bar. = 760 mm.)

203. A tube is filled with mercury and inverted in a trough of mercury. 2 c.cs. of air, measured at the normal pressure, are passed into the tube. The mercury in the tube then stands 44 cm. above the mercury in the trough. Calculate the volume of the air enclosed.

204. A crystal of potassium chlorate gave, on heating, 50 c.cs. of oxygen, measured at 17° C. and 740 mm. Calculate the volume at 0° C. and 760 mm.

205. A piece of marble gave, when dissolved in acid, 240 c.cs. of carbon dioxide, measured at 0° C. and 800 mm. What would be the volume at N.T.P.?

206. 1 grm. of magnesium, dissolved in hydrochloric acid, gave 960 c.cs. of hydrogen, measured at 12° C. and 765 mm. Calculate the volume at 0° C. and 760 mm.

207. 20 c.cs. of oxygen, under a pressure equal to that of 740 mm. of mercury, are subjected to an increase of pressure of 100 cm. of mercury. What change in volume will be produced?

CHAPTER XXIV.

EQUIVALENTS.

We have found from previous experiments that 1 grm. of different elements unite with different amounts of oxygen. Thus

1 grm.	of copper	unites with	0.25	grm. oxygen	(Expt. 77.)
1 „	magnesium	„	0.66	„	(Expt. 69.)
1 „	hydrogen	„	8.00	„	(Expt. p. 71.)
1 „	mercury	„	0.08	„	(Expt. 68.)

If we calculate these results so as to find the weight of the various elements uniting with 1 grm. oxygen we obtain the following.

4	grms.	copper	unite with 1 grm.	oxygen.
1.5	„	magnesium	„	„
0.125	„	hydrogen	„	„
12.5	„	mercury	„	„

We say the amounts of the various elements uniting with the same weight of oxygen are **equivalent** to one another.

When we displaced hydrogen by means of metals we found that to displace the same volume of hydrogen very different amounts of the metals were required. As measured by their power of displacing hydrogen, therefore, we again find that different amounts of metals are equivalent to one another.

Generally, we find that when a number of elements unite with, or displace, the same amount of any one element, the weights of the different elements are all different from one another. It is found that when any element unites with or displaces 1 grm. of hydrogen more than 1 grm. of that element is required. In other words, hydrogen has a smaller equivalent than any other element.

For comparing equivalents, therefore, we take hydrogen as a standard. We say the equivalent of hydrogen is 1. The equivalent of any other element is the amount that will combine with or displace unit weight of hydrogen. As 8 grms. of oxygen unite with 1 grm. of hydrogen the equivalent of oxygen is 8. As 35·5 grms. of chlorine unite with 1 grm. of hydrogen the equivalent of chlorine is 35.5. If we were to determine the weight of the hydrogen we found to be displaced by 1 grm. of magnesium, or zinc, or aluminium, we should then be able to calculate the equivalents of these metals by direct comparison with hydrogen.

Employing the method of Expt. 45, and omitting various corrections, we find that

1 grm. of zinc displaces approximately 360 c.cs. of hydrogen.
1 „ magnesium „ „ 960 „ „
1 „ aluminium „ „ 1300 „ „

Now, 1000 c.cs. of hydrogen weigh 0.09 grm.
Hence

1 grm. of zinc displaces $\frac{0.09}{1000} \times \frac{360}{1} = 0.032$ grm. of hydrogen.

1 „ magnesium „ $\frac{0.09}{1000} \times \frac{960}{1} = 0.086$ „ „

1 „ aluminium „ $\frac{0.09}{1000} \times 1300 = 0.116$ „ „

Or 31.3 grms. of zinc displaces 1 grm. of hydrogen.
11.6 „ magnesium „ 1 „ „
8.6 „ aluminium „ 1 „ „

These figures are, as already mentioned, only approximate, but they serve to illustrate a method whereby the equivalents of some metals may be compared directly with the equivalent of hydrogen, and they also show that the equivalents of different

metals are different from one another. For example, the equivalent of zinc is $2\frac{2}{3}$ as great as that of magnesium.

More correct values for the equivalents of the above-mentioned metals are :

Zinc, - - -	32.5
Magnesium, - -	12.0
Aluminium, - -	9.0

Many substances do not either unite with or displace hydrogen, and so their equivalents cannot be directly determined. In such cases we find the equivalent by determining the amount uniting with or displacing the known equivalent of some other element. Chlorine and oxygen are very often used for this purpose, the equivalent of any substance being determined by finding the amount uniting with 8 parts by weight of oxygen, or 35.5 of chlorine.

From the results quoted at the commencement of this chapter we easily calculate that

8 grms. of oxygen combine with { 1 grm. hydrogen. / 32 „ copper. / 12 „ magnesium. / 100 „ mercury.

Therefore these amounts represent the equivalents of the elements mentioned.

We find, therefore, that

1 grm. of hydrogen unites with 8 grms. of oxygen.

1 „ „ is displaced by 12 grms. of magnesium.

From these figures it would appear that the equivalents of oxygen and magnesium should be as 8 : 12. Now, in Expt. 69 we find that these are just the proportions in which oxygen and magnesium combine with one another.

Many different processes are employed in the determination of various equivalents. A few examples of methods for this determination may be added to those already considered.

EXPT. 190. **To determine the equivalent of tin by converting it into oxide by means of nitric acid.**

Weigh a porcelain crucible and lid ; introduce about 0.5 grm. of tin foil, torn into small pieces, and weigh again. Obtain a glass tube about 8 in. long and drawn out at one end. Dip this

end in strong nitric acid, cover the other end with the finger, and remove the tube. As long as the finger is over the wider end no acid will run out of the tube. Put the drawn-out end into the crucible, keeping it covered as completely as possible with the lid, and by slightly releasing the finger allow a drop of acid to run out on to the tin. When the action that is set up is over, add another drop, and repeat this till all the tin has been

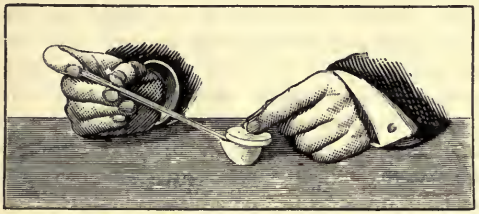

Fig. 85.

acted on. Remove the glass tube, and very carefully heat the crucible till the contents are apparently dry. Then heat as strongly as possible for ten minutes. Allow the crucible to cool, and weigh. Heat for five minutes more, cool, and weigh again.

If the last weighing is very nearly the same as that immediately preceding, it may be regarded as correct. Calculate results as follows :

Second weighing = 17.429 grm. = wt. of crucible + tin.
First „ = 16.741 „ = wt. of crucible.
Difference = 0.688 grm. = wt. of tin.

Third weighing = 17.614 grm. = wt. of crucible + tin oxide.
Second „ = 17.429 „ = wt. of crucible + tin.
Difference = 0.185 grm. = wt. of oxygen with which 0.688 grm. of tin has combined.

Then wt. of tin combined with 8 grm. of oxygen

$$= \frac{0.688 \times 8}{0.185} = 29.7 \text{ grm.}$$

equivalent of tin = 29.7.

EXPT. 191. **To determine the equivalent of copper by first converting it into nitrate and then into oxide.**

We have seen that most nitrates when heated decompose, giving off brown fumes and leaving an oxide of the metal. This reaction may be utilised in the determination of some equivalents. We shall describe the method as applied to the determination of the equivalent of copper.

Weigh a small quantity, 0.3—0.5 grm. of thin sheet copper or copper turnings in a porcelain crucible. Add about two drops of water, and then cautiously add strong nitric acid drop by drop, until all the copper is converted into the nitrate. Dry very carefully, and then heat strongly until nothing is visible, on raising the lid, but black oxide. Allow to cool, and weigh. Calculate results as in the last experiment. It is found that the equivalent obtained by this method is nearly 31.6, the true value.

The relation of copper to oxygen in black copper oxide has now been determined by two separate methods, viz.:

(1) By taking oxygen from the heated oxide by means of coal-gas. (Expt. 77.)

(2) By first making the nitrate and decomposing this by heat, as in the last experiment.

By each of these methods we find that 31.5 grm. copper is combined with 8 grm. oxygen.

The equivalents of magnesium or lead or zinc may be determined in the same way. It is advisable to have the metals in as thin strips as possible in these determinations.

Just as the equivalent weights of metals and of hydrogen may be compared by displacing hydrogen from an acid by the metal, so equivalents of some metals may be compared by displacing one metal from a compound containing it, by another metal. It need scarcely be remarked that if one metal displaces a second from a compound, the second will not be able to displace the first.

EXPT. 192. **To find the weight of zinc required to displace 31·5 grm. of copper from copper sulphate.**

Weigh a small porcelain dish, then put in it about 0.7—1.0 grm. of zinc foil which has been cleaned with emery and cut into small strips, and weigh again.

Half fill the dish with a solution of copper sulphate. Using a glass stirring rod, remove the copper from the surface of the zinc as fast as it is produced. The stirring rod must not be removed from the dish until it is finally washed. It must be short enough, therefore, to remain in the dish by itself. If the solution becomes decolourized, add more copper sulphate solution. When, by feeling about with the end of the glass rod, no more zinc can be discovered, allow the precipitated copper to settle, and pour off the liquid as completely as possible. Wash the copper four or five times, by adding water, stirring, allowing the copper to settle, and then pouring off the water. When the last quantity of water is being added, carefully wash any copper from the stirring rod into the dish. Finally wash the copper twice with small quantities of alcohol, and after thoroughly draining off the alcohol, dry the copper carefully on a sand-bath. The copper must not be made too hot, or part of it will be oxidized. When it is dry allow it to cool, and weigh.

The difference between this weighing and the first gives the weight of copper displaced by the amount of zinc used. Or this amount of copper is *equivalent* to the weight of zinc used.

It is found that 1 grm. of copper requires for its displacement 1.03 grm. of zinc. Or taking the value 31.6 as the equivalent of copper, the equivalent of zinc becomes,

$$31.6 \times 1.03 = 32.55.$$

EXPT. 193. **To find the weight of magnesium required to replace 1 grm. of silver from silver nitrate.**

Weigh very carefully 0.1—0.15 grm. of magnesium in a porcelain dish, as in the last experiment. Pour over this a strong solution of silver nitrate, and find out as before when all the magnesium has disappeared. If the action appears to go on very slowly while some magnesium still remains, it is probable that all the silver has been displaced from the silver nitrate used, and more must be added. The washing and drying are done exactly as in the last experiment. It will be found that about 0.11 grm. of magnesium displaces 1.0 grm. of silver.

In this experiment the first washings may contain considerable quantities of silver nitrate, and should be poured into a silver residue bottle. The precipitated silver should also be preserved.

EXPT. 194. **To find the weight of zinc required to displace 1 grm. of silver from silver nitrate.**

Use 0.3—0.5 grm. of zinc, and proceed exactly as in the last experiment. It will be found that about 0.3 grm. of zinc displaces 1.0 grm. of silver.

Comparing the results of Expts. 193 and 194, we see that:

0.11 grm. of magnesium displaces 1 grm. of silver, and
0.3 „ zinc „ 1 „ „

Or 1.0 „ magnesium „ 9 „ „ and
2.7 „ zinc „ 9 „ „

As measured by their power of displacing silver, therefore, we see that the equivalent of zinc is about $2\frac{7}{11}$ times the equivalent of magnesium, as has been previously found on p. 167.

QUESTION. XXIV.

208. Describe any experiment you have made with the object of determining the equivalent of an element.

209. 0.3697 gram of aluminium liberated 0.04106 gram of hydrogen on being dissolved in a strong solution of sodium hydrate. Calculate the equivalent of aluminium.

210. 0.37 gram of sodium gave, on being converted into the chloride, 0.941 gram. Calculate the equivalent of sodium.

211. 0.2 gram of a certain metal was oxidised by nitric acid. The oxide was found to weigh 0.254 gram. Calculate the equivalent of the metal.

212. What do you understand by the equivalent of an element?

213. 1 gram of magnesium displaces 0.083 gram of hydrogen, 9 grams of silver, and 2.625 grams of copper. What are the equivalents of these elements?

214. Silver chloride is deprived of its chlorine when heated in hydrogen or coal gas. In a certain experiment 0.724 gram of silver chloride lost, when heated in coal gas, 0.179 gram. Calculate the equivalent of silver.

215. 2.0876 grams of hydrogen, when passed over heated copper oxide, yielded 18.7406 grams of water. Calculate the equivalent of oxygen.

216. 5.2632 grams of aluminium when dissolved in caustic potash solution gave off an amount of hydrogen which, when burnt, yielded 5.2562 grams of water. Calculate the equivalent of aluminium.

217. Common salt and hydrogen chloride have the following composition :

	per cent.		per cent.
Sodium, - -	39.77	Hydrogen, - -	2.75
Chlorine, - -	60.23	Chlorine, - -	97.25

Calculate from these figures the combining proportions of chlorine and sodium, taking one part by weight of hydrogen as the unit.

CHAPTER XXV.

LAWS OF CHEMICAL COMBINATION.

It has been found that 1 grm. of magnesium always combines with 0.66 grm. of oxygen to form magnesium oxide, 1 grm. of hydrogen with 8 grm. of oxygen to form water, 100 grm. of mercury unite with 8 grm. of oxygen, 100 grm. of marble will yield 44 grm. of carbon dioxide, and so on.

Facts such as these were observed as soon as the balance came into general use in chemical investigation, and led to the statement of the general rule or law known as the **Law of Definite Proportions.** This law may be stated in the following way :

"The same substance is always composed of the same materials, united in the same proportions."

But it often happens that two or more compounds are known composed of the same elements. Thus carbon monoxide, composed of carbon and oxygen, unites with a further quantity of oxygen to form carbon dioxide. There are also two compounds of copper and oxygen, and no fewer than five compounds of nitrogen and oxygen.

In such cases the relative amounts of the constituents vary, and vary in a perfectly definite manner. Thus, in the black oxide of copper, we have found that the copper and oxygen

stand in the ratio of 1 : 0.254 or 63 : 16. We never find 63 parts by weight of copper combining with 15 or 17 parts by weight of oxygen. In the second compound of these two substances we have found 16 parts of oxygen combining with 126 (2×63) parts of copper. In the case of two of the oxides of lead, we found that, combined with 1 grm. of lead, we have in one case 0.154 grm. of oxygen, and in the other 0.077 grm., and these amounts are in the ratio 2 : 1. Similarly, in the oxides of carbon, we found that in carbon monoxide a certain volume of the gas contains half its volume of oxygen, whereas the carbon dioxide made from it, and therefore containing the same amount of carbon, contains just double the amount of oxygen.

In the case of the five oxides of nitrogen we have, combined with the same weight of nitrogen, weights of oxygen in the ratio of 1 : 2 : 3 : 4 : 5.

These facts illustrate the second law of chemical combination, or **the Law of Multiple Proportions.**

"When two elements unite to form more than one compound, if the amount of one element is kept constant, then the greater amounts of the other element in the various compounds are simple multiples of the least amount of that element."

To account for these laws, Dalton, in 1808, suggested a theory which is generally known as **Dalton's Atomic Theory.**

According to Dalton, all elements are composed of indivisible particles, or **atoms.** The atoms of any one element are all alike in weight and other properties, but differ from the atoms of other elements. Chemical compounds are made by the union of atoms.

Let us see then how Dalton's Atomic Theory would explain the Law of Definite Proportions. Consider the case of black copper oxide. This is made by the union of atoms of copper, all of which are alike in weight, and atoms of oxygen, which also are all alike. Suppose that one atom of copper always unites with one atom of oxygen. Then in any given quantity of copper oxide, there will be equal numbers of copper atoms and of oxygen atoms, and the ratio of the weights of these will be the same as the ratio of the weights of the individual atoms.

Next as to the Law of Multiple Proportions. We have proved that in red copper oxide the amount of copper combined with a certain weight of oxygen is greater than in the case of the black oxide. If we assume that in the black oxide there is one atom of copper combined with one atom of oxygen, then the least further amount of copper that might be induced to combine with the same amount of oxygen would be another atom ; in other words, the amount of copper relative to the oxygen would be doubled. Consequently, combined with any given amount of oxygen, we ought to have twice as much copper as in the black oxide. Expts. 77 and 78 prove that this really is the case. Expts. 79 and 80 illustrate the same point.

If we take a certain amount of any compound, say mercuric oxide, we might divide it into halves, quarters, eighths, sixteenths, etc., and we can imagine that if the division could be carried far enough, we should reach ultimately an amount so small that, on further division, it would cease to be mercuric oxide, and would be resolved into mercury and oxygen.

This smallest portion of mercuric oxide that can exist as such is called a **molecule**, and must obviously contain at least two atoms. Chemists are agreed, for various reasons, that elementary substances themselves do not exist, usually, in the form of free atoms. In most cases the atoms pair off and exist in the form of "double atoms." These double atoms are also called molecules. We may therefore have molecules made of the same kind of atoms, and molecules made of different kinds of atoms.

In general, therefore, we may say that

A **molecule** *is the smallest portion of any substance capable of separate existence.*

An **atom** *is the smallest amount of any element that can enter into the formation of a molecule.*

Atomic weights and molecular weights.

No individual atom or molecule has ever been rendered visible, even by the most powerful microscope, nor has any atom ever been weighed. The relative weights of the atoms can, however, be obtained by finding the proportions in which various elements combine.

Suppose, for example, that an atom of magnesium of weight X unites with an atom of oxygen of weight Y to make magnesium oxide. In a given quantity, say 1 grm. of magnesium, there will be a very great number of atoms, say n, and these will unite with n oxygen atoms. Now, experiment shows that 1 grm. of magnesium unites with 0.66 grm. of oxygen.

Hence we have $$n X = 1$$
$$n Y = .66$$
$$\therefore X : Y = 1 : .66 = 3 : 2.$$

Therefore an atom of magnesium is $1\frac{1}{2}$ times as heavy as an atom of oxygen.

When the atomic weights of the elements are compared, it is seen that the hydrogen atom is lighter than any other. It is therefore taken as the unit for atomic weights. When we say the atomic weight of mercury is 200, we mean an atom of mercury weighs 200 hydrogen atoms. The atomic weight of oxygen is 16, that is, an oxygen atom weighs 16 hydrogen atoms. In the list of elements on p. 179 the figures in the third column represent in all cases the weights of the atoms of the various elements, expressed in terms of the weight of the hydrogen atom.

A molecule weighs as much as the sum of its component atoms. Thus a molecule of hydrochloric acid, which consists of an atom of hydrogen weighing 1, and an atom of chlorine weighing 35.5, weighs 36.5 hydrogen atoms. A molecule of mercuric oxide, which contains 1 mercury atom weighing 200 and 1 oxygen atom weighing 16, weighs 216 hydrogen atoms.

Valency. It might be thought from what has been said about combination taking place between atoms, and from the chapter on equivalents, that an equivalent and an atomic weight are represented by the same number. A moment's consideration will show that this is only true when the elements unite atom for atom. Now this is not always the case.

For example, in Expt 72, we saw that when sodium is put in water, hydrogen is liberated, and sodium takes its place, forming sodium hydrate. Since an atom is the least amount of an element that can enter into the formation of a molecule, each molecule of water that has been acted on, must have lost at least

one atom of hydrogen. Now Expt. 75 shows that if the sodium hydrate that is made acts on aluminium, we get a further quantity of hydrogen, and as aluminium is an element, this hydrogen must have come from the sodium hydrate. Moreover, the amount of hydrogen in the second case is just equal to the amount in the first; hence the hydrogen in each molecule of water acted on must have been divisible by 2; that is, there must have been at least two atoms of hydrogen in the molecule. As the amount of hydrogen in a molecule of water has never been divided further, it is concluded that a molecule of water contains *only* two atoms of hydrogen. No evidence has ever been discovered that the amount of oxygen combined with two atoms of hydrogen in a molecule of water is divisible, and hence it is concluded that the smallest amount of water capable of existing as such contains one atom of oxygen and two atoms of hydrogen. Hence the atomic weight of oxygen is double the equivalent, *i.e.* 16.

It is found that the elements have varying combining values, and that the atom of hydrogen has as low a combining value as any other known element. Some, like chlorine, combine with hydrogen atom for atom. These are called *monads*, or monovalent elements. Others, like oxygen, are such that one atom is able to combine with two atoms of hydrogen. These are called *dyads*, or divalent elements. Some, of which nitrogen is an example, combine with three atoms of hydrogen. These are called *triads*, or trivalent elements. Others, like carbon, are able to combine with four atoms. These are called *tetrads*, or tetravalent elements.

One atom of a tetrad is equal in combining value to two atoms of a dyad. No element is known to combine with five atoms of hydrogen, but many are known such that two atoms are able to combine with five atoms of a dyad, *i.e.* are equal in combining value to ten atoms of hydrogen. Such elements are called *pentads*. The atoms of some elements also combine with three atoms of a dyad. Such elements are called *hexads*.

The combining value of an element, as measured by the number of atoms of hydrogen, or of any other monad, with which one atom can combine, is called its **valency**. The valency of one and the same element may be different in different

compounds. Thus nitrogen in combination with oxygen seems to have all valencies, from one to five.

In the case of a monad element, the equivalent is the same as its atomic weight, in dyad elements the atomic weight is twice the equivalent, in triad elements it is thrice the equivalent, and, generally,

Equivalent × Valency = Atomic weight.

Symbols and Formulae.

In the list of elements on the next page, there appear in the second column **symbols** for each element. In some cases these consist of the first letter of the common name of the element, in others of the first letter with another attached. In a third group the symbol has apparently no connection with the name of the element. In these last cases the symbol is derived from the Latin name; thus copper (cuprum) Cu, sodium (natrium) Na, silver (argentum) Ag, etc. Where the symbol consists of two letters, the first is always a capital, the second a small letter. These symbols are much more than mere shorthand expressions for the elements. In every case a symbol is to be understood to signify *one atom* of the element, and therefore a definite weight of the element, viz. an atomic weight. Thus Hg stands for one atom of mercury weighing 200 hydrogen atoms; Fe, for one atom of iron weighing 56 hydrogen atoms, and so on.

When we represent a compound in symbols, we write the symbols of the constituent atoms side by side. Thus MgO stands for magnesium oxide. HCl stands for hydrochloric acid, KCl for potassium chloride. The symbolic representation of a compound is called a **formula.** But, like the symbols of elements, formulae have a greater importance than that of being merely shorthand representations of the names of substances. Take, for example, the formula for mercuric oxide, HgO. The symbol Hg stands for one atom of mercury; that is the least amount that can enter into the formation of a chemical compound. The symbol O stands for the least amount of oxygen that can enter into the formation of a chemical compound. Hence HgO stands for the least amount of mercuric oxide

capable of separate existence, *i.e.* for the molecule of mercuric oxide. It must also represent a definite weight of mercuric

TABLE OF ELEMENTS.

Name of Element.	Symbol.	Atomic Weight.	Name of Element.	Symbol.	Atomic Weight.
Aluminium,	Al	27	Molybdenum,	Mo	96
Antimony,	Sb	120	Nickel,	Ni	59
Arsenic,	As	75	Niobium,	Nb	94
Barium,	Ba	137	Nitrogen,	N	14
Beryllium,	Be	9	Osmium,	Os	191
Bismuth,	Bi	207.5	Oxygen,	O	16
Boron,	B	11	Palladium,	Pd	106
Bromine,	Br	80	Phosphorus,	P	31
Cadmium,	Cd	112	Platinum,	Pt	195
Caesium,	Cs	133	Potassium,	K	39
Calcium,	Ca	40	Rhodium,	Rh	104
Carbon,	C	12	Rubidium,	Rb	85
Cerium,	Ce	141	Ruthenium,	Ru	103.5
Chlorine,	Cl	35.5	Samarium,	Sm	150
Chromium,	Cr	52	Scandium,	Sc	44
Cobalt,	Co	59	Selenium,	Se	79
Copper,	Cu	63	Silicon,	Si	28
Didymium,	Di	145	Silver,	Ag	108
Erbium,	Er	166	Sodium,	Na	23
Fluorine,	F	19	Strontium,	Sr	87
Gallium,	Ga	70	Sulphur,	S	32
Germanium,	Ge	72	Tantalum,	Ta	182
Gold,	Au	197	Tellurium,	Te	125
Hydrogen,	H	1	Thallium,	Tl	204
Indium,	In	113	Thorium,	Th	232
Iodine,	I	127	Tin,	Sn	118
Iridium,	Ir	192.5	Titanium,	Ti	48
Iron,	Fe	56	Tungsten,	W	184
Lanthanum,	La	138.5	Uranium,	U	240
Lead,	Pb	207	Vanadium,	V	51
Lithium,	Li	7	Ytterbium,	Yb	173
Magnesium,	Mg	24	Yttrium.	Y	90
Manganese,	Mn	55	Zinc,	Zn	65
Mercury,	Hg	200	Zirconium,	Zr	90

oxide. For Hg stands for 200 hydrogen-atoms-weight of mer-mercury, and O stands for 16 hydrogen-atoms-weight of oxygen;

hence HgO represents 216 hydrogen-atoms-weight of mercuric oxide. In the same way every formula stands for one molecule of the particular substance it represents. Any molecular weight is obtained by adding together the weights of the separate atoms constituting the molecule.

We have proved that the least amount of water, *i.e.* a molecule, contains an amount of hydrogen that is divisible into two parts. Hence there are two atoms of hydrogen in the molecule of water; the formula for this is written H_2O. In every case the number of atoms, more than one, of an element occurring in a compound is indicated by a small figure written after the symbol for that element, and below the line. CO_2 represents a molecule of carbon dioxide, and indicates that one atom of carbon is combined with two atoms of oxygen. P_2O_5 represents phosphorus pentoxide, and indicates that two atoms of phosphorus are combined with five atoms of oxygen.

QUESTIONS. XXV.

218. State the Law of Definite Proportions. Describe and explain any two simple experiments which serve to illustrate the Law of Definite Proportions.

219. What do you understand by the equivalent of an element? How is it related to the atomic weight?

220. State the Law of Multiple Proportions. What experiments would you make to illustrate this law?

221. Show how the following facts illustrate the Law of Multiple Proportions: (1) Ferrous sulphide contains 63.6 per cent. of iron and 36.3 per cent. of sulphur, (2) Iron pyrites contain 46.6 per cent. of iron and 53.3 per cent. of sulphur.

222. What do you understand by chemical action? Explain how the recognition of the fact that chemical action always takes place between definite proportions of matter led to the establishment of Dalton's Atomic Theory?

223. What do you understand by the "valency" of an element?

224. It is found that 1 lb. of sulphur unites with 1 lb. of oxygen and also with $1\frac{1}{2}$ lbs. of oxygen. How do you reconcile these facts with the Law of Multiple Proportions?

CHAPTER XXVI.

CHEMICAL EQUATIONS. AVOGADRO'S LAW.

CHEMICAL changes are often represented by symbols. For example, when magnesium is heated in air or in oxygen it burns, uniting with oxygen and forming magnesium oxide. Now, one atom of magnesium unites with one atom of oxygen in this case, and one molecule of magnesium oxide, MgO, is produced, and so we may write

$$Mg + O = MgO.$$

Notice that the signs + and = in this equation have not the significance that they have in Mathematics. The sign + indicates "uniting chemically with," and the sign = indicates "produces."

Magnesium and oxygen are not equal in themselves to magnesium oxide. It is only when they unite that we get magnesium oxide, a substance entirely different in properties from magnesium and from oxygen, and in which the properties of magnesium and oxygen are completely lost.

To represent the decomposition of mercuric oxide by heat, we write

$$HgO = Hg + O.$$

In this case the sign = has the same meaning as before, but the sign +, now on the right-hand side of the equation, has more the meaning that it usually conveys. We may interpret the above by saying mercuric oxide produces mercury *and* oxygen, or mercury *together with* oxygen.

No indication is given in either of the above equations of the conditions under which the chemical change occurs, that heat is required to bring about the decomposition of mercuric oxide, for example, or that heat is produced as a consequence of

the union of oxygen and magnesium. An equation is therefore only an incomplete representation of the action going on.

To express in symbols the action that occurs when zinc is put into sulphuric acid, we have the following :

$$Zn + H_2SO_4 = ZnSO_4 + H_2.$$

The zinc displaces the hydrogen from the acid, and unites with the residue of the sulphuric acid to form the substance we call zinc sulphate.

The decomposition of chalk or marble on heating is expressed thus :

$$CaCO_3 = CaO + CO_2.$$

Here a compound splits up into two other compounds of simpler constitution, calcium oxide and carbon dioxide.

The equations representing the various chemical changes that have been already met with, have been given at the ends of the chapters in the previous pages. They should now be studied, and gradually learnt.

Quantitative significance of Chemical Equations.

Bearing in mind that each symbol stands for a definite weight of an element, viz., the weight of an atom expressed in terms of the weight of a hydrogen atom, we are able to calculate the amounts of various materials taking part in a chemical change. A few examples of such calculations are given below.

Example 1. *Find the weight of oxygen given off on heating* 100 *grms. of mercuric oxide.*

In every question of this kind we must first write the equation expressing the change, and then put under the symbols the weights they represent.

$$\begin{array}{rcl} HgO & = & Hg + O \\ \underbrace{200+16}_{216} & = & 200 \quad 16 \end{array}$$

The molecular weight of mercuric oxide, obtained by adding the separate atomic weights, is equal to the weight of 216 hydrogen atoms. This amount yields 16 hydrogen atoms

weight of oxygen. In any other units the same ratio, 216 : 16 of mercuric oxide to oxygen, is maintained, and so we may say "216 grm. of mercuric oxide yield 16 grm. of oxygen, how much oxygen will 100 grm. of mercuric oxide yield?"

By simple "rule of three,"

	Grm. of Mercuric Oxide.		Grm. of Oxygen.		
As	216	: 100	::	16	: x

$$x = \frac{100 \times 16}{216} = 7.4 \text{ grm. of oxygen.}$$

See how the result of Expt. 68 agrees with this theoretical value.

Example 2. *Find what weight of carbon dioxide is obtained on heating* 100 *grm. of calcium carbonate.*

$$\underset{\underbrace{40+12+48}_{100}}{CaCO_3} = CaO + \underset{\underbrace{12+32}_{44}}{CO_2}$$

Here we see that the molecular weight of calcium carbonate is just 100, and the molecular weight of carbon dioxide is 44. We are not concerned with the weight of the quick-lime, so we need not put in the molecular weight of this substance. We see that 100 grm. of calcium carbonate would give 44 grm. of carbon dioxide. Compare the result of Expt. 123, with this value.

Example 3. *What weight of zinc would need to be dissolved in sulphuric acid to yield* 7.5 *grm. of hydrogen?*

The equation expressing the change is

$$\underset{65}{Zn} + H_2SO_4 = ZnSO_4 + \underset{2}{H_2}.$$

We see that 65 grm. of zinc would yield 2 grm. of hydrogen, hence the answer is obtained in the following way:

	Grm. of Hydrogen.		Grm. of Zinc.		
As	2	: 7.5	::	65	: x

$$x = \frac{7.5 \times 65}{2} = 243.75 \text{ grm. of zinc.}$$

Example 4. *What weight of hydrogen would be theoretically obtainable by the action of steam on* 4 *lbs. of iron nails?*

The equation is

$$\underbrace{3Fe}_{3\times 56} + 4H_2O = Fe_3O_4 + \underset{8}{4H_2}.$$

Hence 168 parts by weight of iron yield 8 parts by weight of hydrogen. For example, 168 lbs. of iron would yield 8 lbs. of hydrogen. The answer is therefore obtained from the following proportion:

	Lbs. of Iron.		Lbs. of Hydrogen.
As	168 : 4	::	8 : x

$$x = \frac{4\times 8}{168} = \frac{4}{21} \text{ lbs. of hydrogen.}$$

Combination of gases by volume.

In 1809 Gay Lussac announced a "law" which may be stated thus:

The volumes in which gases combine bear a simple relation to one another, and to the volume of the resulting product, if that is gaseous.

All the volumes must, of course, be measured at the same temperature and pressure.

A number of examples have already been met with which serve to illustrate this law.

When hydrogen and oxygen were exploded, and the water was kept in the gaseous state, it was seen that

2 *volumes of hydrogen combine with* 1 *volume of oxygen, and* 2 *volumes of gaseous water are produced.*

From experiments described on pp. 143 and 144, we see that

(1) 2 volumes of ammonia, when decomposed, give 4 volumes of the mixed constituents, and

(2) In ammonia, 1 volume of nitrogen is combined with 3 volumes of hydrogen.

Hence

2 *volumes of ammonia contain* 1 *volume of nitrogen and* 3 *volumes of hydrogen.*

On burning carbon in oxygen, it was found that

Any given volume of carbon dioxide contains its own volume of oxygen.

When carbon monoxide was exploded with oxygen, it was found that

2 *volumes of carbon monoxide unite with* 1 *volume of oxygen to make* 2 *volumes of carbon dioxide.*

Similarly, it has been found in every case which has been tested, that the law of Gay Lussac holds.

Avogadro's Law. To account for the law of Gay Lussac, it was suggested by Avogadro in 1811 that in equal volumes of all gases there are the same number of atoms. The word "atom" as used here must be taken to mean sometimes *atom*, sometimes *molecule*, as we now understand these terms—*atom* in the case of an elementary gas, *molecule* in the case of a compound gas.

If we consider that gases combine by weight in the proportion of their atomic weights, or of multiples of these, and we find also that they combine in equal or multiple volumes, the above suggestion seems very probably correct. The fact that gases behave in the same way when subjected to varying temperatures and pressures seems also to support this idea. If we proceed to test the statement by actual examples, however, we find that it cannot be maintained in the form given above.

For we know that

(*a*) 2 volumes of hydrogen unite with 1 volume of oxygen to produce 2 volumes of water.

If we use n oxygen atoms, they must have combined with $2n$ hydrogen atoms (this is known to be actually the case), and $2n$ "atoms," or molecules, as we should now call them, of water should be produced. But according to the atomic theory, the atoms of elements are indivisible, and therefore the smallest particle of water must contain at least an atom of oxygen. Hence we could not form more than n "atoms" of water from n atoms of oxygen.

It is proved experimentally that

(*b*) 1 volume of hydrogen unites with 1 volume of chlorine to form 2 volumes of hydrochloric acid gas.

Here again, n atoms of hydrogen will unite with n atoms of

chlorine, and n "atoms," or molecules, of hydrochloric acid should be produced. But 2 volumes of hydrochloric acid gas will contain twice as many "atoms" as 1 volume of hydrogen, *i.e.* $2n$, and consequently we again arrive at contradictory results.

A further simple assumption made by Avogadro was that the atoms of elementary gases do not exist in the free state, but that they form small groups, generally consisting of two atoms. These small groups are the *molecules* of elementary gases.

Avogadro's Law, as it is now generally termed, may be stated in the following way:

Equal volumes of all gases, at the same temperature and pressure, contain the same numbers of molecules.

Avogadro's Law, stated in this form, would lead us to expect the similar behaviour of gases when subjected to the same variations of temperature and pressure. It also explains completely the Law of Gay Lussac.

Consider first the formation of hydrochloric acid by the union of hydrogen and chlorine. Any given volume of hydrogen will contain a certain number, say n, molecules, or $2n$ atoms. The same volume of chlorine will also contain n molecules or $2n$ atoms. If 1 atom of hydrogen unites with 1 atom of chlorine, to form a molecule of hydrochloric acid we should have altogether $2n$ molecules of hydrochloric acid. This would occupy twice the space occupied by n molecules of hydrogen or of chlorine, and so, as is found by actual experiment, any given volume of hydrogen gives *twice* its volume of hydrochloric acid.

Consider next the union of two volumes of hydrogen with one volume of oxygen to form two volumes of water-vapour. If we have $2n$ molecules of hydrogen, we have only n molecules of oxygen and $2n$ molecules of water. But n molecules of oxygen contain $2n$ atoms, and therefore the formation of a molecule of water is seen to involve the union of an atom of oxygen with two atoms or one molecule of hydrogen.

It is not possible to prove Avogadro's Law directly by experiment. The conclusions derived from it are found to agree so completely with the results of experience, that it is now universally accepted as true.

Determination of molecular weights of gases. When we speak of the densities of gases, we mean generally the *relative* densities. Hydrogen is the lightest gas known, and its density is taken to be 1. The density of any other gas is the number of times any given volume of it is heavier than the same volume of hydrogen. Since in equal volumes of all gases we have the same number of molecules, the densities will be in the same ratio as the weights of the individual molecules, or the molecular weights.

Molecular weights, like atomic weights, are expressed in terms of the weight of an *atom* of hydrogen. Since the molecule contains two atoms, the molecular weight of hydrogen is therefore 2, while the density of hydrogen is 1.

Suppose then that we have a gas whose density is 22. We know that its molecule is 22 times as heavy as a hydrogen molecule, and, therefore, 44 times as heavy as a hydrogen atom. In other words, the number expressing its molecular weight is double the number which expresses its density. A similar remark applies to all other cases, and we may say, therefore,

The molecular weight of a gas is numerically equal to twice its density.

Since equal volumes of all gases contain the same number of molecules, we may say that all molecules occupy the same volume. What this volume is we do not know, but it is convenient to refer to it as **two volumes.** Any molecule of a gas, then, is said to occupy two volumes. As in the elementary gases a molecule consists of two atoms, one atom will occupy **one volume.**

It will be easy now to find the relationship between volumes of gases taking part in chemical changes.

Example 1. *What volume of hydrogen and of nitrogen can be obtained by the decomposition of* 10 *c.cs. of ammonia?*

Writing the equation

$$2NH_3 = N_2 + 3H_2$$

2 mols. 1 mol. 3 mols.

4 vols. 2 vols. 6 vols.

Hence 10 c.cs. of ammonia would yield 5 c.cs. of nitrogen and 15 c.cs. of hydrogen.

Example 2. *What volume of oxygen would be required to burn the hydrogen obtained from* 10 *c.cs. of ammonia?*

From Example 1 we find that 10 c.cs. of ammonia yield 15 c.cs. of hydrogen.

Writing the equation

$$\begin{array}{ccc} 2H_2 & + & O_2 = 2H_2O \\ 2 \text{ mols.} & & 1 \text{ mol.} \\ 4 \text{ vols.} & & 2 \text{ vols.} \end{array}$$

Hence we see the volume of the oxygen will be half the volume of the hydrogen, or $7\frac{1}{2}$ c.cs.

Example 3. *What change in volume occurs on passing* 20 *c.cs. of carbon dioxide over red-hot carbon?*

The equation expressing this reaction is

$$\begin{array}{ccc} CO_2 & + C = & 2CO \\ 1 \text{ mol.} & & 2 \text{ mols.} \\ 2 \text{ vols.} & & 4 \text{ vols.} \end{array}$$

Hence we see the volume of carbon monoxide is twice the volume of the carbon dioxide.

(Notice that the above equations have all been written so as never to represent less than a molecule of any gas. Bearing in mind what has been said previously, this is strictly correct. Nevertheless, equations like $H_2 + O = H_2O$, where single atoms of gases are represented, are often used.)

Relation between weights and volumes of gases.

Questions involving, at the same time, weights and volumes of gases, often arise, and this point must next be dealt with. We know by actual weighing that at normal temperature and pressure

0.09 grm. of hydrogen occupies 1 litre very nearly.

By simple proportion we calculate from this that

2 grm. of hydrogen occupies 22.4 litres very nearly.

That is, one molecular weight of hydrogen, expressed in grms., occupies 22.4 litres.

Now the densities of gases are in proportion to their molecular weights, and so if we calculate the molecular weight of any gas, that number of grams will occupy 22.4 litres. We may say then that

The molecular weight of any gas expressed in grams occupies 22.4 litres.

Since the molecules of elementary gases consist generally of two atoms, we may make for them the alternative statement,

The atomic weight of any elementary gas, expressed in grams, occupies 11.2 litres.

As examples of the application of these rules, we may take the following:

Example 1. *What volume of oxygen will be obtained by the decomposition of* 108 *grm. of mercuric oxide?*

First, write the equation representing the chemical change, then put underneath the atomic and molecular weights, and express these in grams.

Now we know that the atomic weight of oxygen, expressed in grams, occupies 11.2 litres, and this also we write underneath. The equation, therefore, is written thus:

$$\begin{array}{ccc} HgO & = & Hg + O \\ 200+16 & & 16 \text{ grm.} \\ 216 \text{ grm.} & & 11.2 \text{ litres.} \end{array}$$

If 216 grm. of mercuric oxide yields 11.2 litres of oxygen, 108 grm. will yield 5.6 litres of oxygen.

Example 2. *What volume of hydrogen will be obtained by the action of* 50 *grm. of sodium on water?*

$$\begin{array}{cccc} Na & + H_2O = & NaOH + & H \\ 23 \text{ grm.} & & & 1 \text{ grm.} \\ & & & 11.2 \text{ litres.} \end{array}$$

If 23 grm. of sodium yield 11.2 litres of hydrogen, what will 50 grm. yield?

$$\begin{array}{cccccc} & \text{Grm. of Sodium.} & & & \text{Litres of Hydrogen.} & \\ \text{As} & 23 & : 50 & :: & 11.2 & : x \end{array}$$

$$x = \frac{50 \times 11.2}{23} = 24.3 \text{ litres of hydrogen.}$$

Beginners often find some little difficulty in not knowing when to use 2 vols. or 1 vol., and 22.4 litres or 11.2 litres. The difficulty disappears if it is remembered that definite volumes, 11.2 litres or 22.4 litres, refer to definite weights of gases, viz. atomic weights or molecular weights expressed in grams, whereas 1 vol. or 2 vols. are used without regard to weights, but for comparing volumes only of gases. If weights in grams are to be used in calculation, then 22.4 litres or 11.2 litres must be used, in other cases the simplest plan is to use 2 vols. or 1 vol.

QUESTIONS. XXVI.

225. What weight of oxygen is contained in 20 grams of mercuric oxide?

226. What weight of oxygen would be required to burn completely 5 grams of each of the following: Hydrogen, carbon, phosphorus, magnesium?

227. What weight of zinc must be dissolved in sulphuric acid in order to liberate 10 grams of hydrogen? If hydrochloric acid were used instead of sulphuric acid would the amount of zinc required be the same?

228. What weight of oxygen could be obtained by the complete decomposition of 100 grams of pure potassium chlorate? If the chlorate were adulterated with 10 % of chloride, what would be the difference in the amount of oxygen evolved?

229. What weight of pure nitric acid would contain 100 grams of oxygen?

230. Calculate the percentage composition of sulphuric acid.

231. How much zinc must be dissolved in sulphuric acid in order to produce 40 lbs. of crystallised zinc sulphate ($ZnSO_4 . 7H_2O$)?

232. What weight of sodium must be used to liberate 1.2 grams of hydrogen from water? What would be the weight of the sodium hydrate produced?

233. What weight of hydrogen is theoretically obtainable by the action of steam on 1 cwt. of red-hot iron nails?

234. Write equations showing the action of strong sulphuric acid on potassium nitrate, sodium nitrate, and ammonium nitrate. If these three substances were the same price per lb., show which would be the cheapest to use in the preparation of nitric acid.

235. If 126 grams of nitric acid were neutralised by ammonia gas, what weight and properties would the new compound have?

236. Calculate the weight of mercuric oxide required to produce 11.2 litres of oxygen gas, measured at 0° C. and 760 mm. What would be the volume of this amount of gas at 15° C. and 740 mm.?

237. What volume of hydrogen, measured at 10° C. and 730 mm,

would be obtained on dissolving 10 grams of zinc in dilute sulphuric acid?

238. What volume of oxygen would be required for the complete combustion of 5 grams of each of the following: Sulphur, phosphorus, magnesium?

239. 17 c.grms. of sodium are thrown on water. How many c.cs. of hydrogen, measured at 19° C. and 800 mm., would be obtained?

240. What common gases do you know having a density 22? (H = 1).

241. What volume of oxygen would be required to burn 5 grams of carbon to carbon dioxide? What would be the volume of carbon dioxide produced?

242. If 100 c.cs. of carbon dioxide are passed through a long tube containing red-hot carbon, what volume of gas will be produced? Of what will it consist?

243. Calculate the weight of a litre of each of the following gases: Carbon monoxide, nitric oxide, sulphur dioxide, ammonia, marsh gas.

244. 20 c.cs. of carbon monoxide are exploded with 18 c.cs. of oxygen. What volume of gas will be left, and of what will it consist? How would you prove the composition of this residual gas?

245. 5 c.cs. of ammonia are completely decomposed by electric sparks. What volume of gas will be left? 10 c.cs. of oxygen are then added and a spark passed through the mixture, what volume of gas will be left, and of what will it consist?

246. What volume of hydrogen will be contained in 5 grams of water?

247. What volume of carbon dioxide, measured at 15° C. and 730 mm., could be obtained by the action of hydrochloric acid on calcium carbonate?

248. What volume of nitrogen would be left after heating iron wire in 40 c.cs. of nitrous oxide?

249. What weight of sodium hydrate would be required to exactly neutralise 20 grams of hydrochloric acid?

250. Calculate the weight of ammonium nitrate which must be decomposed in order to afford one litre of nitrous oxide, measured at 17° C. and 745 mm.

251. How many volumes of air containing 21 per cent. by volume of oxygen are required for the complete combustion of (*a*) one volume of marsh gas, (*b*) one volume of ethylene? Specify the composition by volume of the products in each case.

252. Calculate the volume of ammonia gas at 10° C. and 740 mm. obtainable from 10 grams of ammonium chloride.

253. State Avogadro's Law, and indicate briefly the evidence on which it is based.

254. What weight of oxygen is required to burn 16 grams of hydrogen to water, and what volume of steam at standard temperature and pressure would this amount of water produce?

255. Calculate the percentage composition of potash alum,

$$K_2SO_4 . Al_2(SO_4)_3 . 24H_2O.$$

256. Coal contains about 2 per cent. of nitrogen. Assuming that 75 per cent. of this amount escapes as ammonia on distillation, calculate the weight of coal required to furnish the sal ammoniac necessary to produce 10 kilograms of *liquor ammoniæ*, containing 50 per cent. of its weight of ammonia.

CHAPTER XXVII.

THE HALOGEN ACIDS.

THE four elements—fluorine, chlorine, bromine, and iodine—generally go by the name of the "halogens," or salt-producers. The last three show the greatest similarity, both in the properties of the elements themselves, and also in the majority of their compounds. In addition to the *similarity* of properties, there is also noticed a gradual *transition* from chlorine to iodine, the element bromine having, in many cases, properties intermediate between those of the other two.

Compounds of chlorine, bromine, and iodine, with metals—called chlorides, bromides, and iodides—occur in sea-water, chlorides being most abundant and iodides least. Common salt, or sodium chloride, is the most abundant compound of chlorine. It is used as the source both of chlorine and of sodium.

Hydrochloric acid, HCl.

Hydrochloric acid is easily made from common salt, or any other chloride, by the action of strong sulphuric acid.

EXPT. 195. Warm a little common salt with strong sulphuric acid in a test-tube. A brisk effervescence is produced, and a colourless gas, hydrochloric acid, is given off.

$$NaCl + H_2SO_4 = NaHSO_4 + HCl.$$

Sodium chloride and sulphuric acid yield sodium hydrogen sulphate and hydrochloric acid.

The acid forms steamy fumes in contact with damp air.

Test the gas with a lighted splint. The light is extinguished, and the gas does not burn. Hold a piece of moistened blue litmus paper in the tube. The colour is changed to red. Bring a glass rod moistened with ammonium hydrate near the mouth of the tube. Dense white fumes are produced. The results obtained in these experiments constitute *tests* for hydrochloric acid.

To collect the gas, the apparatus shown in Fig. 86 is employed. Rock salt in small pieces is put into the flask, and covered with a mixture of strong sulphuric acid and water, in

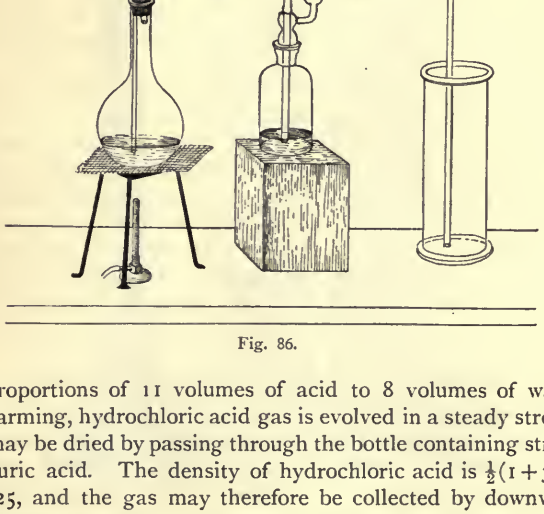

Fig. 86.

the proportions of 11 volumes of acid to 8 volumes of water. On warming, hydrochloric acid gas is evolved in a steady stream, and may be dried by passing through the bottle containing strong sulphuric acid. The density of hydrochloric acid is $\frac{1}{2}(1+35.5) = 18.25$, and the gas may therefore be collected by downward displacement.

Hydrochloric acid is extremely soluble in water being one of

the few gases which can compare with ammonia in this respect. One volume of water dissolves 500 volumes of hydrochloric acid at the ordinary temperature. If a cylinder full of the gas be put mouth downwards in water, the water rapidly rises, showing that the gas is being dissolved. The solubility may also be shown in the same striking way as that of ammonia (p. 141), the water in the trough being coloured with blue litmus. As the gas dissolves the litmus is turned red.

To prepare a solution of hydrochloric acid, the arrangement is similar to that shown in Fig. 86, the gas being washed by bubbling through water contained in the wash-bottle, instead of through sulphuric acid. The solution in water is generally called, simply, hydrochloric acid. An old name for it is "spirit of salt"; sometimes also it is called "muriatic acid." As is always the case with gases, the volume of hydrochloric acid dissolved by water diminishes as the temperature is raised. If a strong solution is heated, hydrochloric acid gas is given off.

EXPT. 196. Heat a little strong hydrochloric acid in a test-tube. Test the gas given off with moist blue litmus paper. It is turned red.

As was seen in Chap. VII., hydrochloric acid attacks many metals, hydrogen being given off.

Like all acids, hydrochloric acid, either gaseous or dissolved in water, neutralizes bases. The formation of sodium chloride by neutralizing sodium hydrate with hydrochloric acid was proved in Expt. 84. We saw in Expt. 195 that in contact with ammonia it formed a white cloud. This consisted of ammonium chloride. The same effect was shown by bringing a cylinder of hydrochloric acid in contact with one containing ammonia (p. 142).

$$NH_3 + HCl = NH_4Cl.$$

The manufacture of Hydrochloric Acid.

In manufacturing hydrochloric acid, salt and sulphuric acid are heated, first in a shallow iron pan *a* (Fig. 87), and afterwards, to complete the action, on the hearths of a furnace, called a reverbatory furnace, *d*, *d*, so arranged that the flame from the fires at *c*, *c* plays directly on the material. In the first stage of the operation, hydrochloric acid and sodium hydrogen sulphate are

produced, but when heated more strongly, the sodium hydrogen

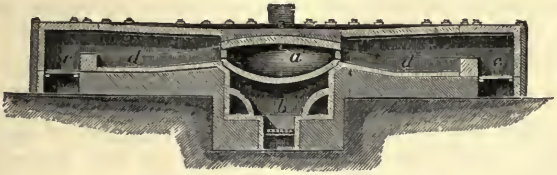

Fig. 87.

sulphate reacts with more sodium chloride, and sodium sulphate is produced.

$$NaCl + H_2SO_4 = NaHSO_4 + HCl,$$
$$NaCl + NaHSO_4 = Na_2SO_4 + HCl.$$

The gas is passed through towers filled with coke over which

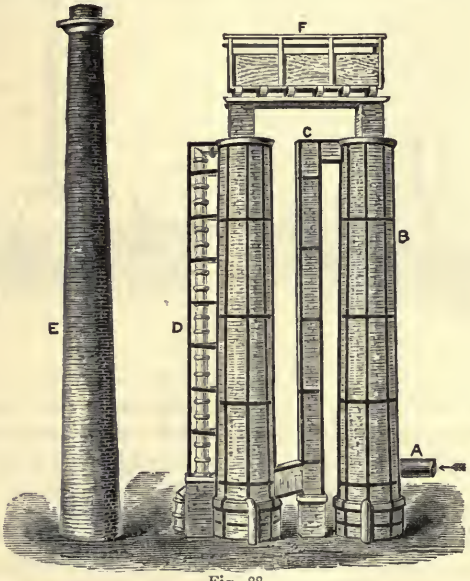

Fig. 88.

water is trickling, the arrangement being as shown in Fig. 88.

The residue obtained in this operation, sodium sulphate, is called "salt-cake," and the whole process is called the "Salt-cake Process." The formation of salt-cake is one step in the production of various sodium compounds from common salt (see p. 213).

To find the relation between the volume of hydrochloric acid gas and the volume of the hydrogen contained in it we may perform the following experiment:

A glass tube, bent into a U-shape, is open at one end, and at the other is provided with a stop-cock (Fig. 89). Near the bend

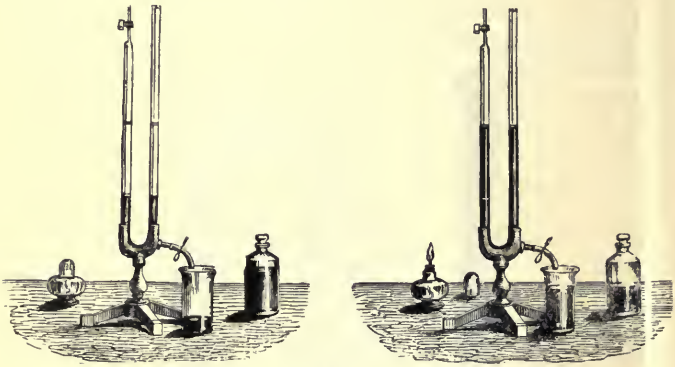

Fig. 89.

a small glass tube is blown on, and this may, when necessary, be closed with a piece of rubber tubing and a clip. The stop-cock is first opened and the clip closed, and the whole apparatus is filled with dry mercury. The stop-cock is now connected to an apparatus delivering dry hydrochloric acid gas, and the clip is opened. Mercury runs out, and hydrochloric acid is drawn into the left-hand limb of the apparatus. The stop-cock is then closed, the mercury is adjusted to the same level in both limbs and the volume of hydrochloric acid gas, under atmospheric pressure, is measured.

The open limb is then filled with mercury mixed with sodium amalgam,[1] corked up tightly, and the gas is transferred several times from one limb to the other. By this means it is brought into contact with the sodium in the amalgam, which unites with the chlorine, and hydrogen is left. This is transferred to the closed limb again, and the volume measured after bringing it to the atmospheric pressure. It is found that hydrochloric acid contains half its volume of hydrogen.

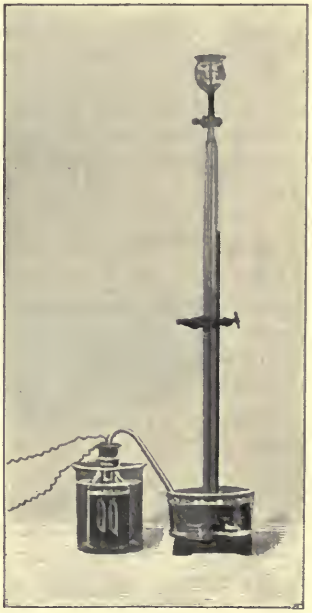

Fig. 90.

The relations between the volumes of hydrogen and chlorine in hydrochloric acid may be shown in the following way:

The bottle (Fig. 90), which is to be kept cool by standing in a vessel of water, contains a strong solution of hydrochloric acid in water. Through the cork pass (1) two platinum wires, to the ends of which pieces of platinum foil are welded, and (2) a glass delivery tube. The platinum wires are connected to a battery of 6 to 10 Grove's cells or their equivalent. The gases produced by the decomposition are allowed to escape until the hydrochloric

[1] To make sodium amalgam, dry mercury is warmed in a wide test-tube, and pieces of dry sodium, about half the size of a pea, are dropped in one at a time. The two metals unite with considerable energy, and each piece of sodium should have united with the mercury before the next piece is added. For the above experiment the amalgam should not be made too pasty. The amalgam acts like diluted sodium. When put into water it generates hydrogen in a steady stream, with none of the violence attending the action of ordinary sodium.

acid solution is saturated with chlorine. An inverted burette, standing over a saturated solution of common salt in water, is then filled with the mixed gases. A strong solution of potassium iodide is then allowed to enter, from a funnel, through the tap. The chlorine reacts with the potassium iodide, liberating iodine, and the liquid is seen to rise in the burette until it is practically half full. The residual gas, on being tested, proves to be hydrogen. Hence we see that of the mixed gases obtained by the electrolysis of hydrochloric acid, half consists of hydrogen and half of chlorine. And we know that hydrochloric acid gas contains half its volume of hydrogen, hence

2 *vols. hydrochloric acid contain* 1 *vol. of hydrogen and* 1 *vol. of chlorine.*

Hydrobromic and Hydriodic Acids.

EXPT. 197. Heat a little potassium bromide with strong sulphuric acid. A gas is given off which fumes in the air as hydrochloric acid did, but in addition a brown gas is produced. In this case we get, not only hydrobromic acid, as we obtained hydrochloric acid, but also, as a result of further action, free bromine.

EXPT. 198. Heat a little potassium iodide with strong sulphuric acid. In this case we get no hydriodic acid given off but only free iodine.

The affinity of hydrogen for bromine is less than its affinity for chlorine, but much greater than its affinity for iodine. When strong sulphuric acid is added to a bromide, the first change results in the formation of hydrobromic acid, but some of this reacts with more sulphuric acid, the hydrogen combining with oxygen and bromine being liberated.

$$H_2SO_4 + 2HBr = 2H_2O + SO_2 + Br_2.$$

A mixture of bromine and hydrobromic acid is therefore obtained.

On adding strong sulphuric acid to an iodide, all the hydriodic acid which is at first produced reacts with the sulphuric acid and so all the iodine is obtained in the free state.

$$H_2SO_4 + 2HI = 2H_2O + SO_2 + I_2.$$

Both hydrobromic acid and hydriodic acid therefore, under these circumstances, play the part of reducing agents.

To prepare hydrobromic acid the following method is employed:

Red phosphorus is placed in a flask and covered with water. The flask is provided with a dropping funnel, containing bromine, and a delivery tube. On allowing the bromine to drop on the mixture, hydrobromic acid is generated, and is

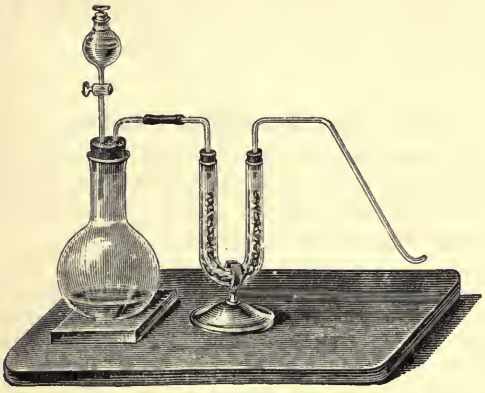

Fig. 91.

freed from bromine vapour by passing through a tube containing moist red phosphorus. The gas is collected by downward displacement.

$$P + 3Br + 3H_2O = H_3PO_3 + 3HBr.$$

A similar arrangement is employed in making hydriodic acid, a mixture of dry red phosphorus and iodine being placed in the flask, and water dropped in from the funnel.

Hydrobromic acid and hydriodic acid are very similar to hydrochloric acid. They are colourless gases, which do not burn, nor allow substances to burn in them. They fume in the air, and form dense white fumes with ammonia. They are readily soluble in water, forming colourless solutions.

The solution of hydrobromic acid is not changed by contact

with air, but the solution of hydriodic acid is rapidly decomposed, the hydrogen being oxidized and the iodine liberated.

$$4HI + O_2 = 2H_2O + I_2.$$

The colour of the solution, therefore, changes to yellow or brown.

It may be shown by a method similar to that employed in the case of hydrochloric acid, that hydrobromic acid and hydriodic acid contain half their volume of hydrogen.

Chlorides, Bromides, Iodides.

The salts formed from hydrochloric acid are called chlorides, from hydrobromic acid, bromides, and from hydriodic acid, iodides. Many of these salts are soluble in water, but a few are insoluble, and often possess characteristic colours.

EXPT. 199. Add silver nitrate solution to solutions containing a chloride, a bromide, and an iodide. In each case a precipitate is produced, the silver chloride being white, the silver bromide pale yellow, and the iodide full yellow.

The changes taking place are shown in the following way:

$$\begin{array}{c|c} Ag & NO_3 \\ Cl & K \end{array} \qquad \begin{array}{c|c} Ag & NO_3 \\ Br & K \end{array} \qquad \begin{array}{c|c} Ag & NO_3 \\ I & K \end{array}$$

Or, writing the equations in the ordinary form:

$$AgNO_3 + KCl = AgCl + KNO_3,$$
$$AgNO_3 + KBr = AgBr + KNO_3,$$
$$AgNO_3 + KI = AgI + KNO_3.$$

These afford good examples of the kind of chemical change known as "double decomposition."

Divide each precipitate into two parts. To one set of three add dilute nitric acid. They are all insoluble. To the other set add dilute ammonia solution. Silver chloride dissolves easily, silver bromide looks as though part had dissolved, the silver iodide is unaffected. To the last two add strong ammonia solution. The bromide dissolves, the iodide does not.

These results may be shown in the following way :

Salt.	Colour.	In Nitric Acid.	In Ammonium Hydrate.
Silver chloride	White	Insoluble	Easily soluble
„ bromide	Pale yellow	Insoluble	Soluble with difficulty
„ iodide	Yellow	Insoluble	Insoluble

Expt. 200. Add lead nitrate solution to solutions of potassium chloride, potassium bromide, and potassium iodide. In each case a precipitate is produced, the lead iodide being yellow, the other two white. Add a little water to each and boil. The precipitates dissolve, but reappear on cooling, the lead iodide appearing in characteristic glistening flakes.

$$Pb(NO_3)_2 + 2KCl = 2KNO_3 + PbCl_2,$$
$$Pb(NO_3)_2 + 2KBr = 2KNO_3 + PbBr_2,$$
$$Pb(NO_3)_2 + 2KI = 2KNO_3 + PbI_2.$$

Expt. 201. Add mercuric chloride solution to solutions of a chloride, bromide, and iodide. No precipitate is obtained in the case of the chloride, or of the bromide, while with the iodide a very characteristic scarlet precipitate is produced. Put a little of this in each of two test-tubes, and add to one more mercuric chloride, and to the other more mercuric iodide. Strangely enough, the precipitate is dissolved by excess of either reagent.

$$HgCl_2 + 2KI = 2KCl + HgI_2.$$

Expt. 202. **To find the weight of silver combined with 1 grm. of chlorine in silver chloride.**

Weigh a small porcelain boat, introduce about 1 grm. of silver chloride, and weigh again. Put the boat with the silver chloride into a hard glass-tube, through which coal-gas may be passed, as in the reduction of copper oxide in Expt. 77. On heating the silver chloride, the chlorine passes off in combination with hydrogen from the coal-gas. After heating for ten minutes,

allow the tube to cool in the stream of coal-gas. Then weigh the boat and its contents, heat for another five minutes in the coal-gas, allow to cool, and weigh again. If the last weighing

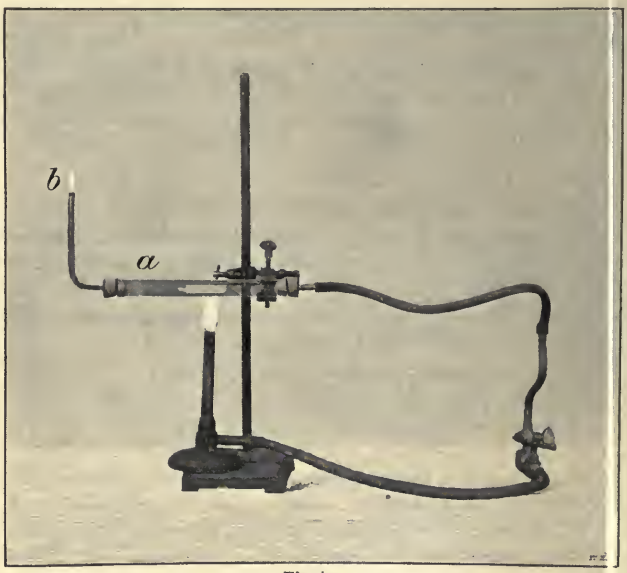

Fig. 92.

differs by only a small amount from the previous one it may be regarded as correct. The following result was obtained in an experiment of this kind:

Second weighing	= 4.704 grm.	= wt. of boat + silver chloride.
First „	= 3.621 „	= wt. of boat.
Difference	= 1.083 grm.	= wt. of silver chloride.
Third weighing	= 4.434 grm.	= wt. of boat + silver.
First „	= 3.621 „	= wt. of boat.
Difference	= 0.813 grm.	= wt. of silver.

1.083 grm. = wt. of silver chloride.
0.813 „ = wt. of silver.
0.270 grm. = wt. of chlorine.

$$\frac{0.813 \times 1}{0.270} = 3.01 \text{ grm.}$$

= weight of silver combined with 1 grm. of chlorine.

The equivalent of silver, or the amount combined with 35.5 of chlorine

$$= 3.01 \times 35.5 = 106.9.$$

For the above experiment the silver chloride may be prepared from silver residues.

QUESTIONS XXVII.

257. Write a summary of the properties of hydrochloric acid.

258. Two bottles are placed before you, each containing a colourless gas. On removing the stoppers, colourless steamy fumes escape from one, and orange-red vapour from the other. What might each of these gases be, and how would you further test them in order to identify them with certainty?

259. Describe with sketch the laboratory method of preparing hydrochloric acid.

260. To decompose a certain amount of salt in making hydrochloric acid in the laboratory, twice as much sulphuric acid is required as is used by the manufacturer in decomposing the same amount of salt. How do you explain this?

261. What happens when you heat (1) hydrochloric acid solution, (2) ammonia solution, (3) a mixture of hydrochloric acid and ammonia solutions?

262. Write equations showing the action of hydrochloric acid on caustic soda, sodium carbonate, black copper oxide. Name the several products.

263. You are required to prepare 100 grams of a solution containing 25 per cent. of hydrobromic acid. Describe the method you would adopt, and calculate the weights of material that would be theoretically requisite.

264. What volume of hydriodic acid could be made from 5 grams of iodine? What weight of chlorine would be required to displace all the iodine from the hydriodic acid?

265. You are required to show that 50 c.cs. of hydrobromic acid contain 25 c.cs. of hydrogen. How would you do this?

266. How would you show that hydrochloric acid contains hydrogen and chlorine?

267. Explain the meaning of the equation:

$$2KI + Cl_2 = 2KCl + I_2.$$

268. Write an account of any three chlorides with which you are acquainted.

269. A gallon of water weighs 70,000 grains. What weight of hydrogen and of oxygen does this amount of water contain? What weight of hydrogen chloride could be produced from the hydrogen?

270. How is hydriodic acid solution prepared? What occurs when it is treated with caustic potash, chlorine, starch paste? In what respect does it differ from hydrochloric acid?

CHAPTER XXVIII.

CHLORINE (Cl.), At. Wt. 35.5. BROMINE (Br.), At. Wt. 80. IODINE (I.), At. Wt. 127.

Chlorine.

EXPT. 203. Heat a little manganese dioxide with hydrochloric acid. A greenish-yellow gas is evolved.

$$MnO_2 + 4HCl = MnCl_2 + Cl_2 + 2H_2O.$$

Test the gas with a lighted splint. The light is extinguished, and the gas does not burn. Test also with a moist litmus paper The colour is destroyed.

This gas is chlorine, a name given to it on account of its greenish colour. It has an extremely disagreeable smell, and even when diluted with air acts violently on the mucous membrane of the nose and throat. On this account, its preparation, except in very small quantities, and experiments with it, must be carried on in the draught cupboard.

The manganese dioxide in the above preparation oxidizes the hydrogen, and then chlorine is liberated. Other oxidizing substances will also liberate chlorine from hydrochloric acid.

EXPT. 204. Heat a little red lead with hydrochloric acid. Chlorine is evolved as before.

EXPT. 205. Heat a mixture of hydrochloric acid with about one third its volume of nitric acid. Chlorine is again evolved. The mixture of hydrochloric acid and nitric acid is called "aqua regia" on account of its power of dissolving the so-called "noble metals," gold and platinum.

$$3HCl + HNO_3 = 2H_2O + Cl_2 + NOCl \text{ (Nitrosyl chloride)}.$$

EXPT. 206. Put a piece of gold leaf into each of two test-tubes. To one add a little hydrochloric acid, and to the other a little nitric acid. In neither case is the metal dissolved. Now mix the contents of the two tubes. The gold is at once dissolved. Platinum is also dissolved by aqua regia. In each case a chloride is made.

Preparation of Chlorine.

If we wish to collect chlorine, we use an apparatus similar to that used in making hydrochloric acid (see Fig. 86). Manganese dioxide in small pieces is heated in a flask with strong hydrochloric acid. It is washed by passing through water, and collected by downward displacement. The density of chlorine is 35.5. Instead of hydrochloric acid, a mixture of common salt and sulphuric acid may be heated with manganese dioxide.

$$MnO_2 + 2NaCl + 2H_2SO_4 = Na_2SO_4 + MnSO_4 + 2H_2O + Cl_2.$$

Chlorine may also be obtained by the electrolysis of a strong solution of hydrochloric acid. The electrodes are generally made of carbon, owing to the fact that chlorine attacks platinum. In the course of an ordinary experiment, however, the extent of this action is not great. The first quantity of chlorine produced dissolves in the water, but this gets saturated before long, and then the chlorine may be collected.

Properties of Chlorine.

Chlorine is an extremely energetic element, uniting directly with many substances. It unites with great energy with antimony, much heat being produced. Thus, when powdered antimony is dropped into chlorine, it falls like a shower of sparks. If leaves of Dutch metal, which consists chiefly of copper, are dropped into a jar of chlorine, they at once take fire. A piece of burning sodium, introduced into a jar of chlorine, continues to burn brightly. Phosphorus takes fire spontaneously when put into chlorine, and burns with an almost non-luminous flame. In all these cases chlorides are produced.

The most striking property of chlorine is its inclination to unite with hydrogen. A jet of burning hydrogen when introduced into a jar of chlorine continues to burn. The colour of the chlorine gradually disappears, and, instead, we get the colourless fuming hydrochloric acid. If chlorine and hydrogen are mixed and exposed to diffuse daylight, they unite slowly; if they are exposed to bright sunlight, or the electric light, or the light of burning magnesium, they unite with an explosion. In the dark, no union takes place.

Chlorine will take hydrogen from many substances containing it. Thus, if a piece of blotting-paper moistened with turpentine, which is a compound of carbon and hydrogen, is put into chlorine, it at once bursts into flame, and much black smoke, carbon, is produced. A lighted candle continues to burn in chlorine, hydrochloric acid being made, and much carbon is deposited.

If chlorine is passed into water, it dissolves, and a solution called **chlorine water** is obtained. A solution saturated at the ordinary temperature contains more than three times its volume of chlorine, and has the colour and odour of the gas. It bleaches litmus readily. On standing, the colour gradually disappears, and hydrochloric acid is found in the solution. The chlorine slowly acts on the water, uniting with hydrogen, and oxygen is set free. If there is something present inclined to react with oxygen, the action goes on more rapidly. This is the explanation of the bleaching action of chlorine.

If chlorine water is cooled to o° C., or if chlorine is passed into ice-cold water, a crystalline solid, called chlorine hydrate ($Cl_2 10H_2O$) is obtained. When this is warmed chlorine is evolved.

Chlorine as a Bleaching Agent.

When a piece of moist litmus paper was put into chlorine, the hydrogen from the water united with the chlorine, and the oxygen acted on the litmus, forming a colourless substance. A piece of damp turkey-red cloth is at once bleached when put into a jar of chlorine. If the chlorine is thoroughly dried by being kept for a few hours in a jar into which some strong sulphuric acid has been poured, and if the cloth be dried before being put into the jar, then nothing beyond a slight change in colour is observed. It appears, therefore, that it is not the chlorine which bleaches, but the oxygen which chlorine is able to liberate from water.

Bromine.

Bromine may be made by the action of oxidizing substances on hydrobromic acid. More generally it is prepared by heating manganese dioxide with a mixture of strong sulphuric acid and a bromide.

$$2KBr + MnO_2 + 2H_2SO_4 = K_2SO_4 + MnSO_4 + 2H_2O + Br_2.$$

Expt. 207. Heat a little manganese dioxide in a test-tube with potassium bromide and strong sulphuric acid. Bromine is given off as a red gas, possessing an exceedingly disagreeable odour. (The word bromine is derived from a word which signifies "a stench.") Test with a lighted splint. The light is extinguished, and the gas does not burn. Hold a piece of moist litmus paper in the tube. The litmus is bleached. To prepare and collect bromine, a mixture of a bromide, manganese dioxide and sulphuric acid, is heated in a retort. The bromine, as it distils off, condenses as a deep red liquid in a cooled flask. The specific gravity of bromine is 3.2. It boils at 59° C. It is soluble to a slight extent in water, 100 grms. of water dissolving rather more than 3 grms. of bromine. The solution has the odour of bromine.

EXPT. 208. Put a piece of litmus paper into bromine water. The colour is destroyed.

If a little bromine be passed down a thistle funnel to the bottom of a glass cylinder, it at once begins to vaporize, and the deep red gas, although much heavier than air, can be seen to be diffusing upwards (Fig. 93). A jet of burning hydrogen introduced into the jar continues to burn, and the colour of the bromine disappears. Instead, we have the colourless fuming hydrobromic acid.

Fig. 93.

Compare this result with that obtained with chlorine.

EXPT. 209. Add a few drops of carbon bisulphide to a little bromine water, shake well, and allow to stand.

The carbon bisulphide forms a separate layer below the water, and is coloured brown by the bromine it has dissolved out from the water.

EXPT. 210. Shake a little ether with bromine water. The ether dissolves out the bromine, and forms a brown layer on the surface of the water.

Iodine.

Iodine is made by distilling a mixture of an iodide and manganese dioxide with strong sulphuric acid.

EXPT. 211. Heat a little potassium iodide with manganese dioxide and strong sulphuric acid. Violet vapours of iodine are given off, and condense as a dark grey solid on the upper parts of the tube.

$$2KI + MnO_2 + 2H_2SO_4 = MnSO_4 + K_2SO_4 + 2H_2O + I_2.$$

EXPT. 212. Put a little iodine in a porcelain dish, cover it with an inverted funnel, and heat gently. The iodine melts, and

very soon changes to a violet vapour. This changes directly to lustrous flat crystals on the cold parts of the funnel. This operation of "subliming" is employed in purifying iodine.

EXPT. 213. Shake a little iodine with water. A faint yellow colour seems to show that a little iodine has been dissolved. Then put in a crystal of potassium iodide and shake again. Iodine is much more soluble in a solution of potassium iodide than in water.

EXPT. 214. Shake a little of the solution with carbon bisulphide. The iodine is dissolved out by the carbon bisulphide, and a violet layer is seen at the bottom of the tube.

EXPT. 215. Shake another portion of the solution with ether. This dissolves the iodine, and carries it to the top, where a brown layer is formed.

EXPT. 216. Add a little starch paste to a dilute solution of iodine. A deep blue solution is obtained. Warm this. The blue colour disappears, but reappears on cooling. This is the best test for iodine.

EXPT. 217. Add a little chlorine water to a solution of a bromide. A yellow colour seems to show that bromine has been liberated. Shake with carbon bisulphide. A brown layer is produced at the bottom of the tube. Bromides are often tested for in this way.

$$2KBr + Cl_2 = 2KCl + Br_2.$$

EXPT. 218. Add a little chlorine water to potassium iodide solution, and shake with carbon bisulphide. A violet solution at the bottom of the tube shows that iodine has been liberated. Iodides are often tested for in this way.

$$2KI + Cl_2 = 2KCl + I_2.$$

EXPT. 219. Add a little bromine water to a solution of potassium iodide, and again shake with carbon bisulphide. A violet solution is again obtained.

$$2KI + Br_2 = 2KBr + I_2.$$

We learn, therefore, that chlorine is able to displace either bromine or iodine from its compounds with metals, and that bromine is able to displace iodine.

EXPT. 220. Add a little starch paste to potassium iodide solution. No result is obtained. Add a little chlorine water. A blue colour is at once produced.

Starch is a test for *free* iodine only.

QUESTIONS. XXVIII.

271. Compare and contrast the properties of chlorine, bromine, and iodine.

272. How would you distinguish a piece of iodine from one of graphite? Describe characteristic properties of each of these substances.

273. In what respects do the chemical properties of chlorine resemble those of bromine?

274. How is the bleaching action of chlorine explained?

275. How could chlorine be obtained from sea-water?

276. Calculate the volume of chlorine gas, measured at N.T.P., obtainable by heating 1 grm. of pure manganese dioxide with excess of hydrochloric acid. (Mn = 55; O = 16; Cl = 35.5.)

277. Describe and explain what happens when (*a*) phosphorus, (*b*) litmus, (*c*) turpentine are introduced into chlorine gas.

278. A given black powder consists of a mixture of iodine and copper oxide. How would you separate and identify these two constituents?

279. Describe any experiment you have seen made with a jar of chlorine, and explain how the jar was filled.

280. Describe any experiments you have made illustrating the properties of iodine. How may this element be obtained from potassium iodide?

281. Two yellow solutions are given to you, one containing free bromine, the other free iodine. What experiments would you make to identify each of these substances?

282. You are provided with a jar of chlorine, how would you make sodium chloride from it? What weight of sodium chloride could be made from (*a*) 1 grm., (*b*) 1 litre of chlorine gas?

283. Turpentine put into chlorine gives rise to a puff of black smoke; phosphorus put into chlorine takes fire; and moist red cloth put into chlorine loses its colour. Explain these changes.

284. By what experiments could you prove that chlorine is a constituent of (1) common salt, (2) potassium chlorate?

285. How is bromine prepared? In what forms does it commonly occur in nature? Give some account of the chief properties of bromine and of the tests by which it may be recognized.

286. A sample of bromine is known to contain a small quantity of chlorine. How would you purify it?

287. A sample of iodine is known to contain a small quantity of chlorine. How would you purify it?

CHAPTER XXIX.

SOME COMPOUNDS OF SODIUM.

THE most abundant compound of sodium is **sodium chloride** or common salt. This occurs in sea-water to the extent of nearly 3 per cent. It is also found in crystalline masses under the earth in Cheshire, Galicia, and in other places. In this latter form it is known as **rock-salt**, and is by no means pure. Sodium chloride crystals are cubic. To obtain the salt from salt-mines, water is poured down and left in contact with the salt. After a time the solution is pumped up, and the water is evaporated. To obtain salt from sea-water the water is evaporated, either spontaneously or by the use of artificial heat, until the sodium chloride crystallizes out.

EXPT. 221. Shake powdered sodium chloride with cold water until no more will dissolve. Then add a little more, and heat the solution. It appears as though no more would dissolve. Pour off the hot liquid from the undissolved portion, and cool it under the tap. Scarcely anything separates out.

This experiment shows that sodium chloride is almost equally soluble in hot and cold water.

EXPT. 222. **To determine the solubility of sodium chloride.**

Use the saturated solution prepared in the last experiment. Weigh a porcelain dish, pour in about 10 c.cs. of the saturated solution, and weigh again. Evaporate off the water as completely as possible on the water-bath, and then dry the salt thoroughly on the sand-bath. Great care is required to avoid loss by spirting at this stage. Allow to cool, and weigh. The calculation of the weight of salt dissolved by 100 grm. of water will be made in the same manner as in Expt. 35. See if your

result agrees with the value shown on the solubility curve on p. 37. It will be noticed that the curve for sodium chloride is almost a horizontal line, indicating the experimental fact that the solubility is scarcely at all increased by raising the temperature.

To obtain pure sodium chloride.

When we wish to get pure sodium chloride we proceed in the following manner:

A saturated solution of common salt in water is made, and filtered from any insoluble impurities. Hydrochloric acid, made

Fig. 94.

in the ordinary way, in the flask *a*, is washed by passing through water contained in the bottle *b*, and then led into the salt solution. The hydrochloric acid is delivered through a funnel which just dips below the surface of the liquid. In this way the risk of the liquid being sucked back owing to the acid being dissolved

too rapidly is avoided. Salt is less soluble in hydrochloric acid solution than in water, consequently, as the acid is dissolved, the salt is precipitated in the form of fine crystals. After washing once or twice with strong hydrochloric acid the crystals are drained, and then heated to drive off the remaining water and hydrochloric acid.

The manufacture of sodium carbonate.

Large quantities of sodium chloride are converted into sodium carbonate, sodium bicarbonate, and caustic soda, by the "Le Blanc Process." The first step in the manufacture of these substances is called the "Salt-cake Process." This has been referred to already under hydrochloric acid. The salt-cake, or sodium sulphate produced in this operation, is heated with coal-dust and limestone in a reverberatory furnace. The sodium sulphate is reduced to sulphide by means of the carbon, and the calcium carbonate decomposes the sulphide, giving sodium carbonate and calcium sulphide.

$$Na_2SO_4 + 4C = 4CO + Na_2S.$$
$$Na_2S + CaCO_3 = CaS + Na_2CO_3.$$

The black mass obtained in this operation is called "Black-ash," and the whole process is called the "Black-ash Process." The black-ash is lixiviated with water, when the sodium carbonate is dissolved out. On evaporating this solution to dryness and heating, a residue called **soda-ash** is obtained. When soda-ash is dissolved in water and the solution is allowed to crystallize, **soda crystals**, or washing soda, is obtained.

EXPT. 223. **To determine the amount of water of crystallization in soda-crystals.**

Weigh about 1 grm. of soda-crystals in a porcelain crucible. Heat gently on a sand-bath till dry (Fig. 95). Then weigh, heat again for 5 mins., and weigh again. See if the result you obtain agrees with the formula $Na_2CO_3 . 10H_2O$.

EXPT. 224. Heat a small quantity of soda-crystals in a dry tube, holding the mouth slightly downwards. Test whether any

carbon dioxide is given off, by pouring the gas into lime-water. No turbidity is produced.

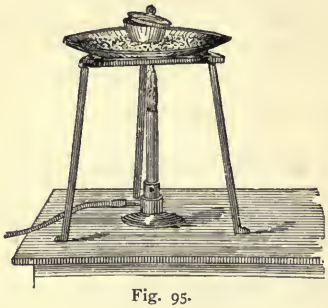

Fig. 95.

Sodium bi-carbonate.

EXPT. 225. Make a saturated solution of soda-crystals in water, and pass carbon dioxide into it. After a while crystals are seen to be separating. Collect these, and dry them. Then heat a small quantity in a dry tube, testing for carbon dioxide as in the previous experiment. Carbon dioxide is readily detected, and at the same time water is seen to be given off. When the residue is cold add a little hydrochloric acid to it. A violent effervescence is set up. Test the gas given off in this case with lime-water. A turbidity is again produced.

The substance obtained on passing carbon dioxide into the solution of sodium carbonate is **sodium bicarbonate**, and it is formed in accordance with the following equation :

$$Na_2CO_3 + H_2O + CO_2 = 2NaHCO_3.$$

In the manufacture of sodium bicarbonate by this method the carbon dioxide is passed over soda-crystals. The water of crystallization in these is more than is required in the reaction represented above.

On heating the dry substance sodium bi-carbonate, or even on heating the solution, it decomposes, giving off carbon dioxide and leaving the normal sodium carbonate.

$$2NaHCO_3 = Na_2CO_3 + H_2O + CO_2.$$

In the **Ammonia-soda Process** a solution of sodium chloride is saturated with ammonia gas, and carbon dioxide is then passed in. Sodium bicarbonate is produced in accordance with the equation

$$NaCl + NH_3 + H_2O + CO_2 = NaHCO_3 + NH_4Cl.$$

As sodium bicarbonate is far less soluble than ammonium chloride, it can be crystallized out.

Sodium hydrate is made by boiling a solution of sodium carbonate with milk of lime.

$$Ca(OH)_2 + Na_2CO_3 = CaCO_3 + 2NaHO.$$

The calcium carbonate is precipitated, and the solution of sodium hydrate is run off and evaporated to dryness in iron pans.

Immense quantities of sodium compounds are employed in the manufacture of glass and of soap.

Preparation of sodium.—Sodium may be obtained from either the carbonate or the hydrate.

1. By heating sodium carbonate to a white heat with carbon,

$$Na_2CO_3 + 2C = 3CO + 2Na.$$

2. By heating sodium hydrate with iron filings,

$$3Fe + 4NaOH = Fe_3O_4 + 4Na + 2H_2.$$

3. By the electrolysis of the fused hydrate. This was the method whereby Davy first prepared sodium.

It has already been observed that sodium is a white shining metal, lighter than water, and easily tarnished. It decomposes water, forming sodium hydrate and liberating hydrogen. It burns in oxygen, forming sodium oxide; and in chlorine, forming sodium chloride. It is soft enough to be easily cut by a knife. On account of the ease with which it oxidizes, it is usually kept under naphtha.

There are only one or two compounds of sodium insoluble in water.

Compounds of sodium are detected by the intense yellow colour they impart to a non-luminous flame.

EXPT. 226. **To find the equivalent of sodium.**

This requires that we should determine the weight of sodium combined with 35.5 grms. of chlorine. Weigh a small dry porcelain dish, then place in it about 0.2 grms. of clean sodium and weigh again as rapidly as possible. Take out the sodium, place a little water—about 5-8 c.cs.—in the dish, and throw in

the sodium a little at a time. Do not put in a second piece until the first has reacted with the water. When all the sodium has been added, neutralize the sodium hydrate solution with hydrochloric acid as in Expt. 84. Then evaporate carefully to dryness, allow to cool, and weigh. Enter results as follows :

Second weighing = 19.427 grm. = wt. of dish + sodium.
First „ = 19.235 „ = wt. of dish.
Difference = 0.192 grm. = wt. of sodium.

Third weighing = 19.720 grm. = wt. of dish + sodium chloride.
Second weighing = 19.427 „ = wt. of dish + sodium.
Difference = 0.293 grm. = wt. of chlorine.

$$\therefore \text{ Equivalent of sodium} = \frac{0.192 \times 35.5}{0.293}.$$
$$= 23.3.$$

Potassium is an element which has the closest possible resemblance to sodium, both in its properties and in its compounds. Almost everything that has been said about sodium applies also to potassium. The elements are obtained by similar processes, and from potassium chloride the sulphate, carbonate, bicarbonate, and hydrate may be obtained, just as the corresponding salts of sodium were obtained from sodium chloride by the Le Blanc process. The ammonia process is not applicable on account of the too great solubility of potassium bicarbonate.

Potassium nitrate is used in the manufacture of gunpowder. It is largely made from sodium nitrate by the action of potassium chloride.

$$NaNO_3 + KCl = KNO_3 + NaCl.$$

Potassium carbonate is found in the ashes of plants.

There are only one or two compounds of potassium insoluble in water. Compounds of potassium may be detected by the violet colour imparted to a non-luminous flame.

QUESTIONS. XXIX.

288. Describe a method of obtaining ordinary table salt from rock salt.

289. What weight of sodium would be contained in 2 oz. of common salt?

290. What effect would be noticed on mixing a strong solution of common salt in water with a strong solution of hydrochloric acid in water? How do you explain the result?

291. Calculate the percentage composition of soda crystals, Na_2CO_3, $10H_2O$.

292. Potassium, sodium, and silver are called monad metals. Why are they so called? How would you distinguish potassium from sodium?

293. Calculate the weight of materials required for the manufacture of 100 kgrm. of caustic soda from crystals of common washing soda. (Na = 23.)

294. What is rock salt? How is it converted into ordinary table salt? Name the metal contained in common salt and describe its properties.

295. How can common salt be converted into soda-ash, soda-crystals, and caustic soda?

296. You are supplied with 500 grams of washing-soda crystals, with the object of making caustic soda therefrom. How would you proceed, and what weight of product would you expect to get?

CHAPTER XXX.

SULPHUR (S), At. Wt. = 32.

SULPHUR, or brimstone, often occurs in nature combined with metals, forming **sulphides.**

Thus we have

Zinc blende, or	zinc sulphide.	ZnS.
Cinnabar,	„ mercury sulphide.	HgS.
Galena,	„ lead sulphide.	PbS.
Iron pyrites,	„ iron disulphide.	FeS_2.

EXPT. 227. Heat a little powdered iron pyrites in a hard-glass test-tube. A gas is given off which condenses on the side of the tube near the top as a yellow powder, and lower down as a yellow or brown liquid. On further cooling, the whole changes

to a yellow solid. If a little of this be burnt, the odour of burnt sulphur will be noticed. The decomposition of the iron pyrites is represented as follows:

$$3FeS_2 = Fe_3S_4 + S_2.$$

Much of the sulphur used is found in volcanic districts. Here it is generally found uncombined, or *native*. To free it from earthy matter it is piled on a sloping surface and fired. The heat produced by burning part of the sulphur, melts the remainder, which flows away. It is afterwards purified by distillation. The arrangement for distilling is shown in Fig. 96. The sulphur is first melted in the iron pot O, and then allowed to run into the retort G, whence the vapour passes into the chamber A. The heat is supplied by a fire placed below G. The first portion of the vapour passing over changes directly to a solid, called flowers of sulphur. Afterwards, when the chamber gets hot, the sulphur forms a liquid, which runs down to the bottom of the condensing chamber whence it is drawn off, and cast into "roll sulphur."

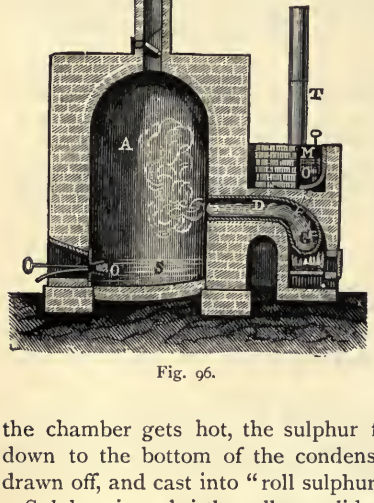

Fig. 96.

Sulphur is a brittle yellow solid which conducts heat very badly. If a piece of sulphur is held in the hand, a crackling sound is heard, and the sulphur often breaks up, owing to unequal expansion.

EXPT. 228. **To observe the behaviour of sulphur when heated.**

Fill a large test-tube to the depth of 1½ in. with coarsely powdered roll sulphur. Heat it cautiously over a bunsen flame, taking the tube out of the flame and shaking fre-

quently when the sulphur begins to melt. If it has not been over-heated, the liquid sulphur will be seen to have a yellow colour. On cooling, *e.g.*, by pouring it into water, it changes again to a brittle yellow solid, which, when broken, is seen to be distinctly crystalline. If the sulphur is further heated, the colour becomes gradually darker, and after a time it is found that the sulphur is becoming thick and viscid. At this stage the colour is almost black. On still further heating the sulphur boils, and changes to a red gas. Near the mouth of the tube this condenses to a yellow powder—flowers of sulphur—while lower down the tube it appears as a yellow or brown liquid.

EXPT. 229. **To make plastic sulphur.**

Pour the sulphur in a thin stream, while near its boiling point, into cold water. A soft, yellow, translucent mass of the consistency of india-rubber is produced. This is called plastic sulphur. Hold a little in the bunsen flame on an iron spoon. It burns with the characteristic blue flame, and the odour of burnt sulphur is noticed.

EXPT. 230. Shake a little plastic sulphur with carbon bisulphide. Pour off a little of the clear solution, and allow it to evaporate. It is found that nothing has dissolved; plastic sulphur is insoluble in carbon bisulphide.

Always allow carbon bisulphide to evaporate spontaneously in the draught cupboard; never attempt to evaporate it by heating.

Sulphur also occurs in two distinct crystalline forms.

EXPT. 231. **To make prismatic sulphur.** Heat a quantity of roll sulphur in a large test-tube until the whole is melted, and allow it to cool slowly. When the part round the side is solid, pour out the still liquid portion, breaking through any crust that may have formed on the surface. Transparent yellow needles are seen projecting inwards from the sides. This is the prismatic form of sulphur.

EXPT. 232. Shake a little of the prismatic sulphur with carbon bisulphide. The crystals dissolve. Allow the carbon bisulphide to evaporate. The crystals which separate are of quite a different shape.

EXPT. 233. **To make the octahedral form of sulphur.**

Powder some roll sulphur, and shake it with carbon bisulphide. It dissolves readily. On allowing the solution to evaporate the third or octahedral form of sulphur is obtained, the crystals being of the shape shown in Fig. 97. This is the most stable form of sulphur, for both plastic sulphur and prismatic sulphur, on keeping, become opaque, changing into a mass of small octahedral crystals. The crystals deposited from the solution of plastic sulphur are also octahedral.

Fig. 97.

These various forms of sulphur differ in density, plastic sulphur having a density of 1.95, prismatic sulphur 1.98, and octahedral sulphur, 2.05.

EXPT. 234. Boil a little sulphur with lime and water. Filter, and to the yellow filtrate add dilute hydrochloric acid. The yellow liquid contains a compound of calcium with a great deal of sulphur, and on decomposing this with hydrochloric acid, calcium chloride and sulphuretted hydrogen are formed, and sulphur is precipitated in the form of a white powder. This is called **milk of sulphur.**

$$CaS_5 + 2HCl = CaCl_2 + H_2S + 2S_2.$$

All the forms of sulphur burn, producing sulphur dioxide which is easily recognized by its odour.

EXPT. 235. **To find the weight of oxygen combining with 1 grm. of sulphur, to make sulphur dioxide.**

Put about 0.25 grm. sulphur into a small porcelain boat, then put this into a hard-glass tube *a*, about 6 in. long, and weigh the whole. The tube *bd* may be made of thin soft-glass tubing about $\frac{5}{8}$ in. bore, and 6 in. long. It is filled for about 4 in., from *c* to *d*, with dried soda-lime, and then from *b* to *c* with pieces of

fused calcium chloride. It is provided at each end with a cork and a short piece of glass tube. This part of the apparatus is also weighed at the commencement of the experiment. Air is drawn through the apparatus by siphoning water from the bottle *f*, and the sulphur is heated until it takes fire. It is impossible to prevent a little sulphur distilling out of the boat into the tube *a*, but care should be taken to have the current of air so slow that no sulphur passes into the tube *bd*. When

Fig. 98.

it is seen that there is no more sulphur in the boat, the apparatus is allowed to cool, and the tubes *a* and *bd*, with their contents, are again weighed. The loss in weight of *a* represents the sulphur burnt. The gain in weight of *bd* represents the weight of sulphur dioxide made. It will be found that the latter is twice the former, so that the sulphur must have combined with its own weight of oxygen.

In making the above experiments different students might use the different forms of sulphur. The result will be found to be the same in each case.

Sulphuretted Hydrogen.

The principal compound of sulphur with hydrogen is sulphuretted hydrogen, H_2S. This is found free in volcanic districts,

and it also occurs in solution in certain mineral waters, *e.g.* in those of the sulphur wells at Harrogate. When organic compounds containing sulphur (such as white of egg), decay, sulphuretted hydrogen is produced. Sulphuretted hydrogen is generally described as smelling like rotten eggs.

EXPT. 236. **To make iron sulphide.**

Make a mixture of iron filings and sulphur in the proportion of 56 parts by weight of iron to 32 parts of sulphur. A greenish powder is produced, which at first sight seems to be a single substance. Shake a little with water in a test-tube, and allow it to settle. Two distinct layers are seen; the iron settles down first, and the sulphur forms a layer, lighter in colour, above this. Put a magnet into some of the powder, and shake it gently. The iron filings are picked out, and the sulphur left. Shake a little of the mixture with carbon bisulphide. Filter through a dry paper, and allow the filtrate to evaporate. Crystals of sulphur are deposited. These experiments show that the green powder is not really a single substance, but consists of iron and sulphur merely mixed together.

Now fill a dry test-tube to the depth of an inch with the mixture, and heat the end of the tube. When the substance becomes red-hot at this point, take the tube out of the flame, and notice that the red glow spreads throughout the mass. When cold, break the tube and examine the residue. It shows no signs of containing two substances. Powder some, and throw it into water. It does not give two separate layers. Neither can iron be picked out with a magnet.

The substance is called iron sulphide, and has been made by the union of iron and sulphur.

$$Fe + S = FeS.$$

By heating copper turnings and sulphur, copper sulphide may easily be made.

EXPT. 237. Pour hydrochloric acid on a little powdered sulphur contained in a test-tube. No change is observed. Then pour hydrochloric acid over iron filings. An inflammable gas, hydrogen, is evolved. Lastly, pour hydrochloric acid over some of the iron sulphide made in the last experiment. An

effervescence is again produced, and the characteristic odour of sulphuretted hydrogen is observed. Hold a piece of paper moistened with lead acetate over the mouth of the tube. A black stain is produced. This is often used as a test for sulphuretted hydrogen.

Apply a light to the mouth of the tube. The gas burns, and the suffocating odour of burnt sulphur is observed.

The equation expressing the formation of sulphuretted hydrogen is as follows :

$$FeS + 2HCl = FeCl_2 + SH_2.$$

If sulphuric acid is used, the equation is

$$FeS + H_2SO_4 = FeSO_4 + SH_2.$$

Owing to its disagreeable odour and its poisonous properties, the gas when required in quantity should always be prepared in the draught chamber.

Preparation of Sulphuretted Hydrogen.

The apparatus required is represented in Fig. 99. The iron

Fig. 99.

sulphide, broken into small pieces, is placed in the bottle *a*.

Dilute sulphuric acid (1 vol. of acid to 7 of water) is poured down the thistle funnel. The gas is washed by passing through water contained in the bottle *b*, and collected by downward displacement. The molecular weight of sulphuretted hydrogen is $(32+2)=34$, so that its density is 17.

A lighted taper is extinguished when put into a jar of the gas, but the gas itself takes fire, and burns with a blue flame. A yellow deposit of sulphur is formed at the same time on the inside of the jar.

The gas is burnt completely when lighted at a jet, water and sulphur dioxide being produced.

$$SH_2+3O=SO_2+H_2O.$$

Sulphuretted hydrogen dissolves in water, forming a colourless solution, which smells strongly of the gas. One volume of water dissolves about three volumes of the gas. The solution has a feebly acid reaction when tested with litmus. Sulphuretted hydrogen is often called hydrosulphuric acid. The solution of sulphuretted hydrogen in water gradually becomes turbid, owing to the hydrogen being oxidized when the sulphur is deposited. Chlorine takes hydrogen from sulphuretted hydrogen very readily. Thus if solutions of chlorine and of sulphuretted hydrogen in water be mixed, a white precipitate of sulphur is at once produced.

The Action of Sulphuretted Hydrogen on Metallic Salts.

EXPT. 238. Make solutions of the following salts acid with hydrochloric acid, and add sulphuretted hydrogen solution to each, or pass the gas directly into the various solutions:

Lead nitrate.	Zinc sulphate.
Copper sulphate.	Cobalt nitrate.
Antimony chloride.	Magnesium sulphate.

In the case of the first three, precipitates are formed, owing to the fact that the sulphides of these metals are insoluble in hydrochloric acid. To the remaining three add ammonium hydrate till the acid has all been neutralized. In the case of the zinc

and cobalt, precipitates appear, while the magnesium sulphate still gives no result.

The following shows the nature of the change which goes on in the formation of the precipitates :

$$\begin{array}{l|l} Cu & SO_4 \\ S & H_2 \end{array} \qquad \begin{array}{l|l} Pb & NO_3 \\ & NO_3 \\ S & H_2 \end{array}.$$

The metal and the hydrogen change places, and metallic sulphides are precipitated.

Many of these sulphides have characteristic colours, and hence sulphuretted hydrogen is used as a test for the metals. Further distinctions depend on the fact that some sulphides are soluble in dilute acids, *e.g.* sulphides of zinc and cobalt, whereas others are not.

The equations showing the formation of the precipitates are as follow :

$Pb(NO_3)_2 + SH_2 = 2HNO_3 + PbS$ (black, insoluble in dilute hydrochloric acid).

$CuSO_4 + SH_2 = H_2SO_4 + CuS$ (black, insoluble in dilute hydrochloric acid).

$2SbCl_3 + 6SH_2 = 6HCl + 2Sb_2S_3$ (orange, insoluble in dilute hydrochloric acid).

The formation of black lead sulphide was used as a test for sulphuretted hydrogen in Expt. 237.

Many sulphides may also be formed by the direct union of the elements. Thus, iron sulphide and copper sulphide have been already made by this method.

The tarnishing of silver is due to the formation of silver sulphide, Ag_2S, by the action of sulphuretted hydrogen occurring in the air. If a drop of sulphuretted hydrogen solution be placed on a silver coin, a dark stain is at once produced.

To Prepare Sulphuretted Hydrogen by the Direct Union of the Elements.

Sulphuretted hydrogen itself may be made by passing hydrogen over boiling sulphur. Hydrogen, generated from zinc and dilute sulphuric acid, is first passed through a bottle *a*,

containing a solution of a lead salt, and then over boiling sulphur contained in the bulb-tube *b*. The sulphuretted hydrogen produced is led into the solution of a lead salt contained in the test-glass *c*, where a black precipitate is produced. In

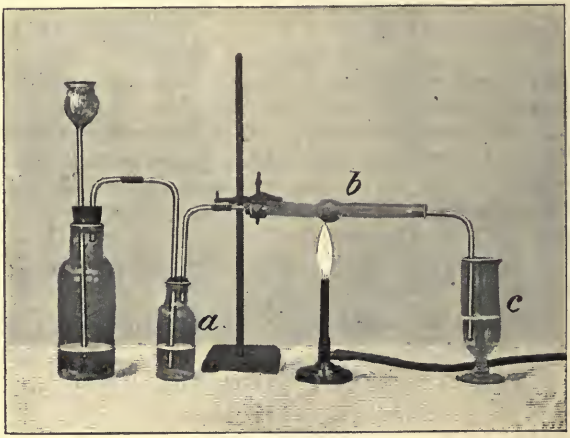

Fig. 100.

many cases a little sulphuretted hydrogen is present in the hydrogen from zinc and sulphuric acid, but unless very impure materials are used, the solution in *a* will never give a black precipitate such as is obtained in *c*.

The Composition of Sulphuretted Hydrogen.

Some metals when heated in sulphuretted hydrogen unite with the sulphur, and hydrogen is liberated. If tin is heated in a known volume of the gas, it is found that the hydrogen left is just equal in volume to the sulphuretted hydrogen used. Hence every molecule of sulphuretted hydrogen contains a molecule of hydrogen, that is, 2 atoms. Now, the density of sulphuretted hydrogen is 17, therefore the molecular weight is 34. And this contains two atoms of hydrogen, weighing 2, therefore the

amount of sulphur in a molecule of sulphuretted hydrogen is 32. This we know to be the atomic weight of sulphur, hence a molecule of sulphuretted hydrogen contains 2 atoms of hydrogen and 1 atom of sulphur, or its formula is H_2S.

Kipp's Apparatus for making Sulphuretted Hydrogen.

As sulphuretted hydrogen is often required in the laboratory, it is advisable to have a ready means of preparing the gas. A Kipp's apparatus is most convenient for this purpose. This consists essentially of two parts, one, the lower, consisting of two glass globes *b* and *d*, joined by a narrow neck *a*, the upper consisting of a single globe provided with a long stem, which is ground into the opening of the lower part at *c*, and passes loosely through the narrow neck *a*. A short tube provided with a tap fits airtight into the globe *b* at *e*. Pieces of iron sulphide are placed in the globe *b*, and, the tap at *e* being open, acid is poured through the funnel until the globe *d* is filled, and the acid comes in contact with the iron sulphide. Sulphuretted hydrogen is evolved, and, passing out through *e*, is washed by the water in the bottle. When the tap is closed, the sulphuretted hydrogen which is still generated, fills *b* and partially fills *d*, forcing the acid out of contact with the iron sulphide into the upper globe. The evolution of gas will thus be stopped until the tap is again turned, and the acid allowed to enter *b*.

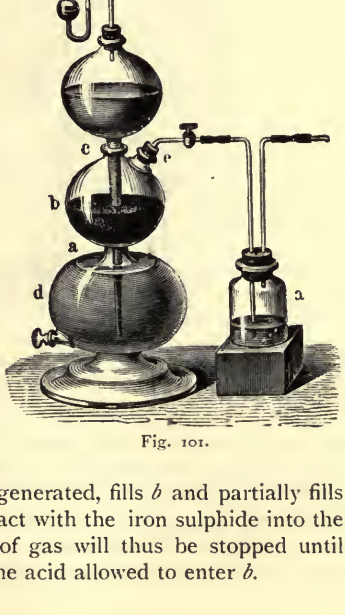

Fig. 101.

A Kipp's apparatus may also be usefully employed for the preparation of carbon dioxide or hydrogen when considerable quantities of these are required.

EXPT. 239. **To find the weight of sulphur combined with 1 grm. of copper in one of the copper sulphides.**

There are two sulphides of copper. To find the weight of sulphur united with 1 grm. of copper in one of these sulphides, proceed as follows :

Weigh a porcelain crucible and lid. Put about 1 grm. of copper turnings in the crucible and weigh again. Now add an

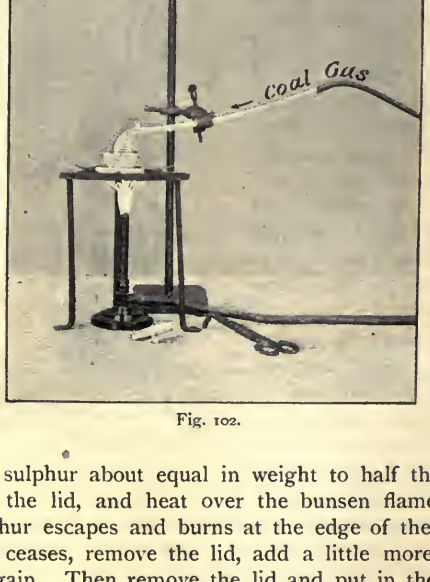

Fig. 102.

amount of sulphur about equal in weight to half the copper cover with the lid, and heat over the bunsen flame. Some of the sulphur escapes and burns at the edge of the crucible When this ceases, remove the lid, add a little more sulphur and heat again. Then remove the lid and put in the crucible the bowl of a tobacco-pipe through which coal-gas is passing

After a few minutes, remove the bunsen, but allow the coal-gas to pass until the crucible is cold. This ensures that only one sulphide of copper shall be formed. Weigh the crucible, and by difference obtain the weight of sulphur which has combined with the copper used. Calculate from your result which of the two sulphides of copper, Cu_2S or CuS, you have made.

QUESTIONS. XXX.

297. By what tests could you prove that a substance said to be sulphur actually was that element?

298. Some purchased milk of sulphur is suspected to be adulterated with plaster of Paris. How could you test it, and by a simple process ascertain the amount of the impurity?

299. How is sulphur commonly obtained? Give an account of the action of heat on sulphur.

300. Describe the methods of obtaining the crystalline forms of sulphur.

301. Sulphur is heated until it is just melted, and part of it is then cooled quickly by pouring it into water. The remainder is heated until it begins to boil, and this is then poured into water. What difference in the result will be noticed in the two cases?

302. Explain why sulphur disappears when burnt in air. What becomes of it? How could you arrange an experiment to show that there is really no loss or gain in weight when sulphur burns in air?

303. A given powder is composed of 56 grams of iron and 32 grams of sulphur. How would you ascertain whether the powder was a mechanical mixture or a chemical compound of the two elements?

304. Write a summary of the properties of sulphuretted hydrogen.

305. How would you show that sulphuretted hydrogen contains (*a*) sulphur, (*b*) hydrogen?

306. How is sulphuretted hydrogen generally prepared? What are the products of its complete combustion in air? How would you separate the element sulphur from the gas? What weight of sulphur is contained in 100 grams of sulphuretted hydrogen?

307. Describe an experiment to show the solubility in water of sulphuretted hydrogen. Why does the solution become turbid on standing?

308. Write equations which show what happens when sulphuretted hydrogen is (*a*) mixed with chlorine, (*b*) heated with tin, (*c*) led through an aqueous solution of sulphurous acid.

309. Required 1 litre of sulphuretted hydrogen gas at 0° C. and 760 mm. pressure; calculate the weight of iron sulphide necessary for its production.

CHAPTER XXXI.

OXIDES OF SULPHUR.

THE two most important oxides of sulphur are:

Sulphur dioxide, SO_2.
Sulphur trioxide, SO_3.

Sulphur Dioxide.

Sulphur dioxide is made when sulphur burns in air or in oxygen.

$$S + O_2 = SO_2.$$

It is also made when many sulphides of metals are heated in air or in oxygen.

EXPT. 240. Heat a little powdered iron pyrites in a narrow glass tube, open at both ends, and held in a slanting position. By this arrangement a current of air is made to pass up the tube, and the sulphur is oxidized to sulphur dioxide, which is readily recognized by its smell.

$$2FeS_2 + 11O = Fe_2O_3 + 4SO_2.$$

When required in quantity in the laboratory, sulphur dioxide is usually prepared by heating a metal, usually copper, with strong sulphuric acid.

EXPT. 241. Cover a few copper turnings in a test-tube with dilute sulphuric acid, and warm. No action takes place. Pour off the acid, add a little strong sulphuric acid, and warm again. Soon a brisk effervescence is set up, and the odour of sulphur dioxide is perceptible. At the same time the copper becomes coated with a black substance. The action that goes on is not thoroughly understood, but it is approximately represented as involving (1) the formation of copper sulphate and hydrogen,

and (2) the reduction by the hydrogen of another molecule of sulphuric acid, with the formation of sulphur dioxide and water.

1. $Cu + H_2SO_4 = CuSO_4 + H_2$.
2. $H_2 + H_2SO_4 = H_2O + H_2SO_3$ $(H_2O + SO_2)$

or $Cu + 2H_2SO_4 = CuSO_4 + 2H_2O + SO_2$.

EXPT. 242. Heat a little carbon with strong sulphuric acid. Sulphur dioxide is again produced.

$$C + 2H_2SO_4 = CO_2 + 2H_2O + 2SO_2.$$

Preparation of Sulphur Dioxide.

For collecting the gas—this operation should be carried on in a draught cupboard—the apparatus used in the preparation

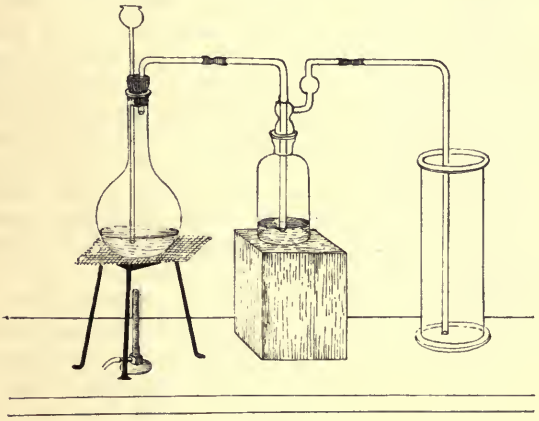

Fig. 103.

of chlorine and hydrochloric acid is employed. Copper turnings are put in the flask provided with a thistle funnel, down which strong sulphuric acid is poured. On heating, the gas is readily obtained in a steady steam, and after bubbling through water contained in the wash-bottle is then collected by downward displacement. The density of sulphur dioxide is $\frac{1}{2}(32 + 32) = 32$.

Properties of Sulphur Dioxide.

Sulphur dioxide is a colourless gas, with a pungent suffocating smell. It extinguishes a light, and does not itself burn. If a jar is inverted over water, the water rises owing to the solution of the gas. The gas is easily liquefied by the application of either cold or pressure. A convenient method for liquefying sulphur dioxide is to pass the gas, after bubbling through strong sulphuric acid, into a long narrow tube surrounded by a mixture of ice and salt, as shown in Fig. 104. The liquid is colourless, and boils at $-10°$ C., giving off large volumes of sulphur dioxide gas. The warmth of the hand is sufficient to cause it to boil vigorously. If a little water be poured into the tube, the sulphur dioxide boils away rapidly, and at the same time the temperature of the water falls below 0° C, and ice is formed.

Fig. 104.

Where a number of demonstrations have to be made it will be found convenient to use a siphon of liquid sulphur dioxide instead of preparing the gas in the usual way.

The apparatus for the preparation of a solution of sulphur dioxide in water is the same as that shown in Fig. 103, the collecting cylinder being replaced by a vessel of water. One volume of water at the ordinary temperature dissolves about 45 volumes of the gas. The solution contains an acid called sulphurous acid, H_2SO_3.

$$H_2O + SO_2 = H_2SO_3.$$

This is very unstable, and on warming readily gives off sulphur dioxide.

Sulphurous acid and sulphur dioxide readily unite with oxygen.

On this account they act as reducing agents. A solution of sulphur dioxide in contact with air gradually loses its odour, and the solution is found to contain sulphuric acid. They will also take oxygen from various coloured substances, converting them into colourless bodies. Thus sulphur dioxide is used for bleaching substances that would be injured by bleaching powder, *e.g.* sponges, blankets, etc.

EXPT. 243. Add sulphurous acid to an infusion of red rose leaves.[1] A pale straw-coloured liquid is obtained.

In certain cases sulphur dioxide acts as an oxidizing substance.

EXPT. 244. Mix solutions of sulphur dioxide and sulphuretted hydrogen. A precipitate of sulphur is at once obtained.

$$SO_2 + 2H_2S = 3S + 2H_2O.$$

Here sulphur dioxide supplies oxygen, which combines with the hydrogen in the sulphuretted hydrogen.

The same result is obtained if two jars, one of which contains sulphur dioxide and the other sulphuretted hydrogen, are brought mouth to mouth. The sulphur forms a yellow deposit on the sides of the jars. If the gases have been previously dried, the liberation of the sulphur in this experiment takes place very slowly. It has been suggested that the sulphur found in volcanic regions has been produced in this way.

Composition of Sulphur Dioxide by Weight.

The composition of sulphur dioxide by weight was shown in Expt. 235. It was there seen that 1 grm. of sulphur unites with 1 grm. of oxygen.

Volumetric Composition of Sulphur Dioxide.

The volumetric composition of sulphur dioxide is shown in the same way as that of carbon dioxide. On burning sulphur in a measured volume of oxygen, it is found that the volume of sulphur dioxide made is the same as that of the oxygen used.

[1] To make the infusion of rose leaves, boiling water is poured on to dried red rose leaves. The liquid is filtered, and the filtrate is made acid with a little dilute sulphuric acid.

Sulphites.

When sulphurous acid is neutralized by bases, salts called sulphites are produced.

EXPT. 245. **To prepare sodium sulphite.**

Put about half a test-tubeful of sulphurous acid into a porcelain dish, and add sodium hydrate solution, a little at a time. Stir after each addition, and test a drop of the solution on a glass rod with litmus paper. Presently it will be found that the litmus is not affected. The odour of sulphur dioxide will at the same time be found to have disappeared. Evaporate the solution to a small bulk and allow to cool. Crystals of sodium sulphite separate out.

$$2NaOH + H_2SO_3 = Na_2SO_3 + 2H_2O.$$

Drain these, and dry them with blotting paper.

EXPT. 246. To separate quantities of sodium sulphite add (*a*) dilute sulphuric acid, (*b*) strong sulphuric acid. Why is there no effervescence in the first case, whereas abundance of sulphur dioxide is obtained in the second? On warming the dilute acid sulphur dioxide is readily obtained.

$$Na_2SO_3 + H_2SO_4 = Na_2SO_4 + H_2O + SO_2.$$

EXPT. 247. Make a solution of sodium sulphite, and add to it barium chloride solution. A white precipitate of barium sulphite is obtained.

$$BaCl_2 + Na_2SO_3 = BaSO_3 + 2NaCl.$$

Add dilute hydrochloric acid. The precipitate is dissolved. Now boil this solution with strong nitric acid, or add chlorine water. The barium sulphite is oxidized to barium sulphate, and this, being insoluble in dilute acid, appears as a white precipitate.

Sulphur Trioxide.

When sulphur is burnt in air or in oxygen, the gas produced generally appears rather misty. This is due to the simultaneous formation of a slight amount of sulphur trioxide, SO_3.

Sulphur trioxide is obtained in much greater quantity by using something which, in a sense, coaxes the oxygen and sulphur

dioxide to combine. Finely divided platinum, generally called platinum black, is very effective for this purpose. For making sulphur trioxide, the platinum is conveniently used deposited on asbestos. A little asbestos fibre is saturated with platinum chloride solution, and then heated in the blow-pipe flame. The chloride is decomposed, and the platinum is left on the asbestos. A solution of sulphur dioxide in water is placed in a Wolff's bottle, and air is blown through. Sulphur dioxide will be carried

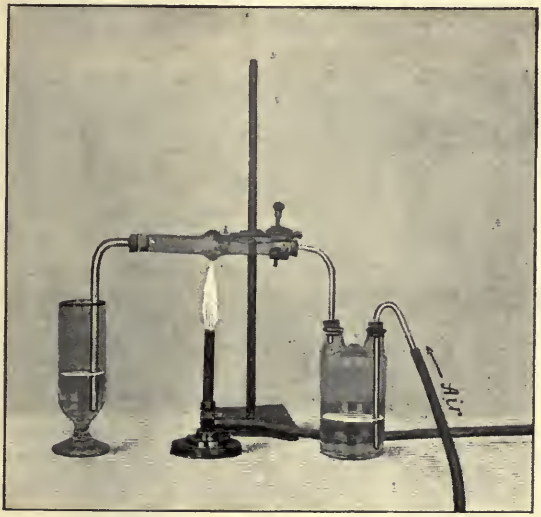

Fig. 105.

along by the escaping air, and the mixture passes through a tube, in the middle of which is a bulb containing the platinized asbestos. As long as the platinum is cold, a colourless gas escapes from the tube. On heating it, however, it determines the union of sulphur dioxide with oxygen from the air, and a white cloud of sulphur trioxide escapes into the air. On passing this into water, a solution of sulphuric acid is obtained.

$$SO_3 + H_2O = H_2SO_4.$$

Sulphur trioxide is also obtained by heating "fuming" sul

phuric acid. The acid is heated in a retort, and the receiver is cooled by a mixture of ice and salt. Sulphur trioxide collects as white silky-looking crystals.

$$H_2S_2O_7 = H_2SO_4 + SO_3.$$

On dropping a little into water, the two unite with a hissing noise, sulphuric acid being produced.

Sulphuric Acid.

Sulphuric acid is manufactured by oxidizing sulphur dioxide. The process may be imitated in the following way:

Fig. 106.

Into a large glass flask *d* are passed (Fig. 106)

1. **Sulphur dioxide,** made by burning sulphur in the bulb tube *b*.

2. **Nitric acid,** made from potassium nitrate and sulphuric acid in the retort *c*.

3. **Steam** from the flask *a*, which contains boiling water.

4. **Air**, over and above that used by the burning sulphur, blown through the bulb-tube *b*.

The changes which go on are approximately as follow: The sulphur dioxide is first oxidized by the nitric acid, nitrogen peroxide being formed at the same time.

$$SO_2 + 2HNO_3 = H_2SO_4 + 2NO_2.$$

If the supply of nitric acid is stopped, and more sulphur dioxide is passed, the brown colour of the nitrogen peroxide gradually disappears, and nitric oxide is formed.

$$SO_2 + NO_2 = SO_3 + NO.$$

The sulphur trioxide unites with water, giving sulphuric acid.

$$SO_3 + H_2O = H_2SO_4.$$

The nitric oxide unites with oxygen from the air, again forming NO_2.

$$NO + O = NO_2.$$

This can oxidize more sulphur trioxide, and so the oxidation is, essentially, carried on by the oxygen of the air. The nitrogen peroxide acts as a carrier of oxygen, and theoretically a small amount of nitric acid ought to suffice for the preparation of an indefinite quantity of sulphuric acid.

On the large scale the glass flask is replaced by a large leaden chamber. The sulphur dioxide is obtained by roasting pyrites, and the nitric acid by the action of sulphuric acid on sodium nitrate. The gases escaping from the chamber, which consist largely of nitrogen from the air, pass up a tower, the "Gay Lussac Tower," filled with coke, over which sulphuric acid trickles. This dissolves the oxides of nitrogen present in the escaping gas, and is then pumped to the top of another tower, the "Glover Tower." As it trickles down the Glover tower it meets the hot gases from the pyrites burners, which carry the oxides of nitrogen along with them back again to the chamber. The acid is not allowed to get above a certain strength—about 70 per cent. of sulphuric acid—in the chamber. It is afterwards concentrated by boiling off the water, first in leaden pans, and afterwards in either glass or platinum vessels. It is purified by distillation.

Sulphuric acid, often called oil of vitriol, is a heavy oily liquid of specific gravity 1.8. It boils at 338° C. It has a great affinity for water, hence it is often used for drying gases. Its use for this purpose has been frequently referred to in the previous pages.

When it is mixed with water, much heat is evolved. The mixing should be done very carefully, the acid being poured in a thin stream into the water. When the commercial acid is diluted, a white precipitate of lead sulphate is obtained. This has been formed by the action of the sulphuric acid on the leaden chamber. It is soluble in the strong acid, but insoluble in water or dilute acid.

In preparing carbon monoxide from formic or oxalic acid, the function of the sulphuric acid was to separate the elements of water (see p. 121).

Many organic substances are charred by the action of sulphuric acid.

EXPT. 248. Write on paper with dilute sulphuric acid. The writing is invisible until the paper is warmed. When the water has evaporated the sulphuric acid chars the paper, and the writing is rendered visible.

Sulphates.

The salts of sulphuric acid are called sulphates. They may be prepared either by dissolving metals in sulphuric acid, or by neutralizing the acid with metallic oxides. When metals are dissolved we get hydrogen evolved when the acid is cold and dilute (see p. 40), and sulphur dioxide if the acid is hot and strong (see p. 230).

Crystals of copper sulphate have been already prepared by the action of sulphuric acid on copper oxide (see p. 85). Other sulphates would be prepared in a precisely similar manner.

Many sulphates occur in nature, *e.g.*:

Epsom salts, magnesium sulphate, $MgSO_4 . 7H_2O$.
Barytes or heavy spar, barium sulphate, $BaSO_4$.
Gypsum and selenite, calcium sulphate, $CaSO_4 . 2H_2O$.

Nearly all sulphates are soluble in water. Barium sulphate and lead sulphate are insoluble. Calcium sulphate and strontium sulphate are slightly soluble.

EXPT. 249. Add a solution of barium chloride to a solution of sodium sulphate, or to dilute sulphuric acid. A white precipitate of barium sulphate is produced.

$$BaCl_2 + H_2SO_4 = BaSO_4 + 2HCl.$$

Add dilute hydrochloric acid. The precipitate is unaffected.

EXPT. 250. Add a solution of lead nitrate to dilute sulphuric acid. A white precipitate of lead sulphate is produced.

$$Pb(NO_3)_2 + H_2SO_4 = PbSO_4 + 2HNO_3.$$

Sulphates may be deprived of their oxygen and reduced to sulphides.

EXPT. 251. Heat a little sodium sulphate, mixed with sodium carbonate, on charcoal in the reducing flame. Place the residue on a silver coin, and moisten with a drop of water. A black stain of silver sulphide shows that the sodium sulphate has been converted into a sulphide. Or the residue may be heated with dilute hydrochloric acid, and a piece of paper moistened with lead acetate held over the mouth of the tube. A black stain of lead sulphide will be obtained.

The above are the tests chiefly relied on for detecting soluble sulphates.

To distinguish a sulphite from a sulphate :

1. Sulphates do not give sulphur dioxide when acted on by an acid, as we have seen sulphites do.

2. They may be distinguished by their barium salts. Barium sulphate, as we saw in Expt. 249, is insoluble in hydrochloric acid ; barium sulphite is soluble.

QUESTIONS. XXXI.

310. How is sulphur dioxide prepared from (*a*) sulphur, (*b*) sulphuric acid, (*c*) a sulphite? Write an account of its properties.

311. What is formed when sulphur dioxide is dissolved in water? Describe an experiment to show the solubility in water of sulphur dioxide.

312. How would you distinguish between sulphur dioxide and carbon dioxide?

313. Describe a method of proving that sulphur dioxide contains its own volume of oxygen.

314. How is sulphur trioxide obtained from sulphur dioxide, and sulphur dioxide from sulphur trioxide?

315. Write an account of the properties of sulphur trioxide.

316. Compare and contrast the bleaching action of chlorine with the bleaching action of sulphur dioxide.

317. Describe the action of the oxides of sulphur on solutions of carbonate of soda.

318. What are the properties common to acids? In what respects does sulphurous acid differ from sulphuric acid?

319. How is sulphur dioxide made from iron pyrites?

320. How is sulphur dioxide liquefied, and what are the properties of the liquid?

321. Distinguish by chemical formulæ between a sulphide, a sulphite, and a sulphate. What is the action of sulphuric acid on (*a*) iron sulphide, (*b*) sodium sulphite?

322. How can sulphuric acid be converted into sulphurous acid, and sulphurous acid into sulphuric acid? How could you distinguish a sulphite from a sulphate?

323. Calculate the weight of sulphur required to make 100 tons of sulphuric acid (S = 32). Briefly explain the method by which sulphuric acid is made.

324. How many volumes of oxygen are contained in 100 volumes of sulphur dioxide? How many grams of sulphur are contained in 100 grams of sulphur dioxide?

325. How is sulphur dioxide usually made on a small scale? How is it made by the vitriol maker?

326. What volume of sulphur dioxide, measured at normal temperature and pressure, is theoretically obtainable by burning 1000 grams of iron pyrites containing 48 per cent. of sulphur?

327. What is the product of burning sulphur in (*a*) oxygen, (*b*) hydrogen?

328. What weight of sulphur is contained in 10 litres of sulphur dioxide gas?

CHAPTER XXXII.

VOLUMETRIC. ACIDIMETRY AND ALKALIMETRY.

We shall now try some further experiments on the neutralization of acids and alkalies. We must first become acquainted with other measuring vessels, besides the burette and the measuring cylinder which we have already used.

(1) A **Pipette** enables us to measure definite volumes of a liquid with great ease. The most useful form for our purpose is one that will *deliver* 20 or 25 c.cs. of liquid. The form shown in Fig. 107 is very convenient. The lower end is narrowed considerably, but, as usually bought, not nearly enough. The opening may be made smaller without appreciably affecting the capacity of the pipette, by holding the end just in contact with a bunsen flame for a little while. Care must be taken not to seal up the end, or even to make the opening *too* small.

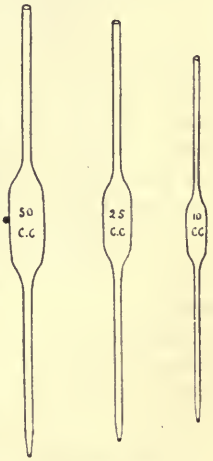

Fig. 107.

A ring is marked on the upper part of the pipette, and after being filled to this mark with liquid, the pipette is supposed to deliver a volume equal to that indicated on the glass.

To use the pipette, put the narrow end below the surface of a liquid, and by sucking at the other end, draw the liquid in until the surface is above the mark on the stem. Then quickly cover the upper end with the first finger, and remove the pipette from the liquid. If the end is tightly closed, no liquid will escape, but by gradually releasing the finger, the

liquid may be allowed to run out slowly. As in the case of the burette, the surface of the liquid is not plane. The usual plan is to arrange so that the bottom part of the curve is on a level with the mark on the pipette. When the liquid is allowed to run out, it will be found that after the continuous stream ceases, a little more gradually drains down and collects at the bottom of the pipette. If, after allowing the liquid to drain for say 10 seconds, the surface of the liquid which has been received in a flask or other vessel, is touched with the end of the pipette, some more will be drawn out, but not the whole. It is important, in order that results may be strictly comparable, that the same amount of liquid should be left in the pipette each time, and therefore the length of time left for draining should be the same in every experiment.

(2) **Measuring flasks** are flasks which, when filled to a mark on the neck, contain a definite volume of liquid; those holding 250, 500, or 1000 c.cs. are most generally used. The capacity is marked on the flask.

Indicators enable us to tell whether a liquid is acid, alkaline, or neutral. We have had to use *litmus* for this purpose on many occasions. Another useful indicator is *methyl orange*.[1] The solution of this becomes pink when made acid, and yellow when alkaline.

A **standard solution** is one which contains a known weight of any substance in a known volume of the solution.

EXPT. 252. **To prepare a standard solution of oxalic acid.**

Weigh a watch-glass, place on it about 5 grms. of oxalic acid, and weigh again. Transfer this to a 250 c.c. flask in the following way: Put in the neck of the flask a funnel about 3 in. diameter, the leg of which has been cut down to about $\frac{1}{4}$ in. Hold the

[1] This is an artificial colouring-matter of complex constitution.

To prepare the methyl orange solution, dissolve 1 grm. of methyl orange in alcohol, then make up to 1 litre with dilute alcohol. Solutions must not be hot when methyl orange is used as the indicator.

Methyl orange cannot be used as the indicator when organic acids, *e.g.* oxalic acid, are being used.

watch-glass over the funnel, gradually tilt it, and, using a wash-bottle, slowly wash the oxalic acid through the funnel into the flask. If the funnel should get choked up, keep the watch-glass over the funnel, but do not wash off any more of the oxalic acid till that in the funnel has been washed through.

When the whole of the acid has been introduced into the flask, and the watch-glass and funnel have been washed thoroughly, the flask should not be much more than half full

Fig. 108.

of liquid. By shaking this round, dissolve all the acid, then fill up to the mark with water, and mix thoroughly by shaking. Put the value of the solution, *e.g.*,

5.2 grm. oxalic acid in 250 c.cs.
0.0208 grm. oxalic acid in 1 c.c.,

on a label on the flask.

From a standard solution, any given weight of the dissolved substance may readily be taken.

Suppose 5.2 grm. have been made up to 250 c.cs. of solution,

then each 25 c.cs. taken out will contain 0.52 grm. of oxalic acid, 100 c.cs. would contain 2.08 grm., and so on.

It must be continually borne in mind that every part of the work, no matter how insignificant it may appear, must receive strict attention if the best results only are to be obtained. It would, for example, require a very awkward student to fail to introduce, say 90 per cent. of any given solid into a flask, but very great care must be exercised if quite the whole of what has been weighed is to be so introduced.

EXPT. 253. **To prepare a standard solution of sodium carbonate.**

Proceed in exactly the same way as in Expt. 252. The sodium carbonate should have been thoroughly dried by heating in a porcelain dish, and allowed to cool before weighing. From 4 to 5 grm. of sodium carbonate in 250 c.cs. will be sufficient. Label the solution as before.

EXPT. 254. **To find the weight of sodium carbonate required to neutralize 1 grm. of crystallized oxalic acid.**

Carefully wash out the burette with water, then rinse twice with a few c.cs. of the oxalic acid solution. In each case let the acid run away through the tap. Then fill up the burette until the level of the liquid is above the zero mark. Allow the liquid to run out slowly, taking care that the little tube below the tap is filled with the solution, until the surface is opposite to or below the zero mark. Measure out 25 c.cs. of the sodium carbonate solution into a 4 oz. flask, add a few drops of litmus solution—not sufficient to colour the liquid very deeply—and commence to run in the acid about 2 c.cs. at a time. Carbon dioxide will be produced, and some of this will remain in solution. Continue to add the acid until the colour of the litmus changes to red. The carbon dioxide in solution will help to redden the litmus, and so the liquid must be boiled to expel this. If the colour changes to blue on boiling, add more acid until the litmus is permanently reddened. The first experiment is in the nature of a trial. Suppose the litmus were blue when 26 c.cs. of acid had been added, and red when 28 had been added. Then the true value will be somewhere between these.

Take a fresh quantity of the sodium carbonate solution, add as much litmus as before, take the burette reading, and run in at once 26 c.cs. of the acid. Then boil, and add more acid (a drop at a time) till the colour of the litmus is purple after boiling. Read off the volume of acid used. Repeat two or three times to confirm your result. If the volumes required in various experiments differ by 0.1 or 0.2 c.c., the average may be taken as the true value. If there is considerable difference between the results obtained, it will be evident that the experiments are not being properly made. A fresh series should in this case be commenced.

Methyl orange is not affected by carbon dioxide, and so, when it is possible to use it, it is the most convenient indicator when a carbonate is to be tested. It cannot be employed, however, if oxalic acid is being used.

Enter results in the following manner:

Acid Solution.

5.2 grm. crystallized oxalic acid in 250 c.cs.
0.0208 „ „ „ 1 c.c.

Sodium Carbonate Solution.

25 c.cs. used in each experiment.
4.3 grm. sodium carbonate in 250 c.cs.
0.43 „ „ „ 25 c.cs.

Acid in Burette. Litmus Indicator.

Expt.	Burette Readings.		c.cs. Used.
	1st.	2nd.	
Trial—			Between 24 and 26.5—
1	21.7	46.3	24.6
2	0	24.5	24.5
3	24.5	49.1	24.6
			Mean = 24.6 c.cs.

Wt. of sodium carbonate in 25 c.cs. of solution = 0.43 grm.

This neutralizes 24.6 c.cs. oxalic acid solution.

24.6 oxalic acid solution contain 24.6×0.0208 grm. oxalic acid. Then, if 0.43 grm. sodium carbonate neutralizes (24.6×0.0208) grm. oxalic acid, what weight of sodium carbonate will neutralize 1 grm.?

	Grm. Oxalic Acid.			Grm. Sodium Carbonate.
As	(24.6×0.0208)	: 1	::	0.43 : x

whence $$x = \left(0.43 \times \frac{1}{24.6 \times 0.0208}\right) \text{grm.} = 0.84 \text{ grm.}$$

EXPT. 255. **To find the strength of the hydrochloric acid supplied, using a standard sodium carbonate solution.**

We might use the standard carbonate solution to determine the weight of acid, say hydrochloric acid, in any volume of a given solution. The acid should be diluted until a convenient volume—say between 10 and 25 c.cs.—is required to neutralize 25 c.cs. of the carbonate solution. If the volume of acid required is much below 10 c.cs., an error of 0.1 c.c. will represent a considerable fraction of the whole. If more than 25 c.cs. of acid are needed, much time is spent in filling up the burette. Whenever it is found by trial that a very small volume is required from the burette, take 50 or 75 c.cs. of the carbonate solution instead of 25 c.cs., or else make up 250 c.cs. of the acid solution to some greater volume—say 500 c.cs. or 1000 c.cs. with water.

The **reaction** is thus represented:

$$\underset{106}{Na_2CO_3} + \underbrace{2HCl}_{2 \times 36.5} = 2NaCl + H_2O + CO_2$$

$$73$$

Hence 73 grm. hydrochloric acid neutralize 106 grm. sodium carbonate.

Method. After preparing the burette as in the last experiment, fill it up with the acid. Put into two separate flasks about 40 or 50 c.cs. of water, and one drop of methyl orange. Put into one of these flasks a drop of the acid solution, and into the

other a drop of the carbonate solution. Then in performing the experiment the object will be to obtain in the experimental flask a colour intermediate between the colours in the comparison flasks. Take 25 c.cs. of the sodium carbonate solution for each experiment. Put into this one drop of the methyl orange solution.

Methyl orange is not affected by carbon dioxide, hence there is no necessity for boiling. In fact methyl orange should never be used with hot solutions.

Enter results in the following manner:

Sodium Carbonate Solution.

25 c.cs. used in each experiment.
4.8 grm. sodium carbonate in 250 c.cs.
∴ 0.48 „ „ „ 25 c.cs.

Acid in Burette. Methyl Orange Indicator.

Expt.	Burette Readings.		C.Cs. Used.
	1st.	2nd.	
Trial—			Between 17 and 18.8—
1	18.8	36.2	17.4
2	0.0	17.3	17.3
3	17.3	34.6	17.3
			Mean = 17.3 c.cs.

Calculation. 17.3 c.cs. of the acid solution neutralize 0.48 grm. sodium carbonate. From the equation we see that

106 grms. sodium carbonate are neutralised by 73 grms. acid.
Hence

1 „ „ „ is „ $\frac{73}{106}$ „

and

0.48 „ „ „ „ „ $\frac{73 \times 0.48}{106}$ „

Now this weight of acid is contained in 17.3 c.cs. of the solution
∴ In 1 c.c. of solution there would be

$$\frac{73}{106} \times \frac{0.48}{1} \times \frac{1}{17.3} \text{ grm. acid} = 0.0191 \text{ grm. hydrochloric acid.}$$

In 1 litre there would be 19.1 grm. hydrochloric acid.

The strengths of solutions of nitric acid, sulphuric acid, or sulphurous acid, might be determined in the same way.

The operation of testing the strength of a solution by using another standard solution is known as **titration.**

Having standardized the hydrochloric acid solution, this might be used for standardizing a solution of an alkali. Sodium hydrate and potassium hydrate cannot be obtained in the pure state, and so the exact strength of a solution of each of these has to be determined by this roundabout process.

EXPT. 256. **To find the strength of a solution of sodium hydrate, using the standard hydrochloric acid solution.**

The **reaction** is thus represented :

$$\underset{40}{NaOH} + \underset{36.5}{HCl} = NaCl + H_2O.$$

Hence 40 grm. of sodium hydrate neutralize 36.5 grm. hydrochloric acid.

Method. Put the acid in the burette as before. Use 25 c.cs. of sodium hydrate for each experiment. Either methyl orange or litmus may be used as indicator. It might be worth while using litmus for part of the experiments, and methyl orange for the remainder. As no carbon dioxide is evolved, no boiling is required.

Enter results thus :

Acid Solution.

0.0191 grm. acid per c.c. solution.

Expt.	Burette Readings.		C.Cs. Used.	Indicator.
	1st.	2nd.		
Trial—			Between 22 and 24—	
1	24.0	47.1	23.1	Litmus.
2	0.0	23.3	23.3	„
3	23.3	46.5	23.2	Methyl orange
4	0.0	23.2	23.2	„
			Mean = 23.2 c.cs.	

Calculation. 23.2 c.cs. of acid solution contain
0.0191 × 23.2 grm. hydrochloric acid = 0.443 grm. hydrochloric acid.
36.5 grm. hydrochloric acid neutralize 40 grm. sodium hydrate.

∴ 1 „ „ neutralizes $\frac{40}{36.5}$ „

and

0.443 „ „ neutralizes $\frac{40 \times 0.443}{36.5}$ „

But the sodium hydrate neutralized was contained in 25 c.cs. of solution.

∴ 25 c.cs. solution contain $\frac{40 \times 0.443}{36.5}$ grm. sodium hydrate,

and 1 c.c. solution contains $\frac{40 \times 0.443}{36.5} \times \frac{1}{25}$ grm. sodium hydrate

$= 0.0194$ grm.

∴ 1 litre solution contains 19.4 grm. sodium hydrate.

Expt. 257. **To find the percentage of potassium hydrate** in ordinary stick potash, **using a standard hydrochloric solution.**

The **reaction** is

$$KOH + HCl = KCl + H_2O.$$
$$56 \quad 36.5$$

Hence we see that 56 grm. potassium hydrate neutralize 36.5 grm. hydrochloric acid.

Method. Put a stick of caustic potash in a dry test-tube provided with a cork, and weigh. Allow the potash to slide out of the tube down the neck of a 250 c.c. flask. Cork up the tube, and weigh again. The difference is the weight of the stick potash taken. Dissolve the potash, taking care to wash the neck of the flask, make the solution up to 250 c.cs., and mix thoroughly.

Enter results in the following way:

First Weighing = 17.401 grm. = Weight of tube + potash
Second „ = 11.520 „ = Weight of tube.
Difference = 5.881 grm. = Weight of potash.

This was made up to 250 c.cs. of solution.

∴ 25 c.cs. solution contain 0.5881 grm. potash.

Acid Solution.

1 c.c. = 0.0191 grm. HCl.

Acid in Burette. Methyl Orange Indicator.

Expt.	Burette Readings.		C.Cs. Used.
	1st.	2nd.	
Trial—			Between 16 and 17.8—
1	17.8	34.9	17.1
2	0.0	17.1	17.1
3	17.1	34.3	17.2
			Mean = 17.1 c.cs.

Calculation. 17.1 c.cs. acid solution contain

$$17.1 \times 0.0191 = 0.3266 \text{ grm. hydrochloric acid.}$$

To neutralize this we require

$$\frac{56}{36.5} \times \frac{0.3266}{1} \text{ grm. potassium hydrate.}$$

And this weight must be contained in 0.5881 grm. of the specimen used.

Hence we find the percentage of pure potassium hydrate thus:

$$\text{As} \quad 0.5881 : 100 :: \left(\frac{56}{36.5} \times 0.3266\right) : x$$

$$x = \frac{100}{0.5881} \times \left(\frac{56}{36.5} \times 0.3266\right)$$

$$= 85.2.$$

Hence of the stick potash 85.2 per cent. is pure potassium hydrate.

The percentage of real sodium hydrate in stick soda, or of sodium carbonate in soda crystals may be determined. The method will be precisely similar to that of the last experiment.

QUESTIONS. XXXII.

329. 300 c.cs. of a solution of caustic potash containing 40 grams KHO per litre are required to neutralise 100 c.cs. of sulphuric acid, and the same volume of nitric acid. Calculate the amount of acid in each solution.

330. Describe the method by which you prepared any standard solution. For what purpose did you use the solution?

331. 25 c.cs. of a solution of sodium hydrate, containing 40 grams NaOH per litre, were neutralised by 19.7 c.cs. of a certain sulphuric acid solution. What was the strength of the sulphuric acid solution?

332. By what method would you ascertain the amount of acetic acid in a given sample of vinegar?

333. Calculate the percentage of sodium hydrate in a sample of commercial soda from the following data:

5 grams commercial soda made up to 250 c.cs. solution.
25 c.cs. solution neutralized by 20.7 c.cs. sulphuric acid solution.
1 c.c. acid solution contains .0245 gram H_2SO_4.

334. 25 c.cs. of a solution of sodium hydrate, containing 40 grams NaOH per litre were neutralised by 21.2 c.cs. of a solution of sulphurous acid. What was the weight of SO_2 in 1 litre of the solution?

335. 25 c.cs. of a given sodium carbonate solution neutralized 19.1 c.cs. of a sulphuric acid solution. The acid solution contained 49 grams H_2SO_4 per litre. What was the weight of Na_2CO_3 in a litre of the given solution?

336. A normal solution of sulphuric acid contains 49 grams of H_2SO_4 in 1 litre. What weight of crystallized sodium carbonate, $Na_2CO_3\,10H_2O$, would be exactly neutralized by 50 c.c. of this solution? (Na = 23.)

MISCELLANEOUS QUESTIONS.

337. Write equations showing the action of hydrochloric acid on caustic soda, sodium carbonate, and ammonia gas. Name the several products.

338. What weight of oxygen is required to burn 16 grams of hydrogen to water, and what bulk of steam at standard temperature and pressure would this amount of water produce?

339. Why was oxygen gas so called? What objection might be urged against the name?

340. Two bottles are given to you, one of which contains bromine and the other nitrogen peroxide. Describe how you would proceed to distinguish between these substances.

341. How do you explain that the atmosphere has nearly the same composition all over the world?

342. Express in words all that is implied in the following equations:

$$NH_3 + HCl = NH_4Cl,$$
$$2CO + O_2 = 2CO_2.$$
$$ZnO + H_2SO_4 = ZnSO_4 + H_2O.$$

343. Name any bodies formed by the action of sulphuric acid, concentrated and dilute, hot and cold, upon the following substances: Sodium chloride, zinc, copper, nitre, and oxalic acid. Give equations in each case.

344. What action has chlorine at the ordinary temperature on the following: carbon, hydrogen, phosphorous, water, sulphuretted hydrogen, solution of sulphurous acid?

345. Explain the meaning of the syllables *ous* and *ic* in such names as *nitrous* acid and *nitric* acid, *sulphurous* acid and *sulphuric* acid.

346. Nitric acid is poured upon copper, hydrochloric acid upon chalk, and sulphuric acid into water. What are the changes which occur in each case, and the appearances which accompany each reaction?

347. What is meant by the term "allotropic modification"? Give examples of some allotropic modifications of non-metallic elements.

348. Explain the following terms: analysis, synthesis, mixture, compound, acid, neutral, alkaline.

349. By what simple tests or experiments could you distinguish and identify nitric, hydrochloric, and sulphuric acids?

350. Given water, an iron tube, a furnace for heating the same, and some charcoal, what gases could be prepared, and how?

351. Which of the following gases burn in air, and in which will the flame of a taper be extinguished: carbon dioxide, nitrous oxide, nitric oxide, carbonic oxide, sulphuretted hydrogen, marsh gas, oxygen, nitrogen, ammonia, hydrochloric acid?

352. What volume of carbon dioxide, at 0° C. and 760 mm. can be obtained from 10 grams of pure marble? What changes occur when the pure gas is passed (*a*) into lime-water, (*b*) over heated metallic sodium, (*c*) over red-hot charcoal?

353. How would you prove that 2 volumes of steam contain 2 volumes of hydrogen?

354. Why is gunpowder considered to be a mechanical mixture of saltpetre, charcoal, and sulphur, and not a chemical compound of these constituents?

355. If hydrogen sulphide is passed into a mixture of iodine and water, sulphur is deposited, and hydrogen iodide formed; the iodine may be in turn displaced from combination by bromine, and the bromine by chlorine. What weight of each of these elements would be required to displace the sulphur contained in 100 grams of hydrogen sulphide?

$$S = 32,\ I = 127,\ Br = 80,\ Cl = 35.5.$$

356. Give illustrations of (1) simple combination between two elements, (2) simple decomposition of a chemical compound, (3) the union of two compounds, (4) the formation of two compounds by double decomposition.

357. Explain the meaning in Chemistry of the word "molecule." Write in words the meaning of each of the following symbols:

$$H,\ O,\ O_2,\ 3H_2,\ NH_3.$$

358. A solution contains either carbonate of sodium, chloride of sodium, or sulphate of sodium. How would you ascertain which of the three was present?

359. What is the effect of heating a crystal of copper sulphate? Metallic copper may easily be obtained from the aqueous solution of the salt. How?

360. If air contains 23 per cent. of its weight of oxygen, how many lbs. of carbon must be burnt in order to remove all the oxygen from 500 lbs. weight of air?

361. Three samples of water are given to you. One is a hard water, another a soft water, and the third is sea-water. What chemical experiments would enable you to discriminate between them?

362. Describe experiments by which you would establish completely the formula NH_3 for ammonia.

363. How would you prepare *pure* sulphuretted hydrogen gas? What means would you adopt to test the validity of the expression H_2S for this compound?

364. If 100 parts by weight of air contains 23 of oxygen and 77 of nitrogen, how many tons of air would be needed to burn 1 ton of coal consisting of

Carbon,	90.55
Hydrogen,	4.14
Nitrogen,	1.26
Oxygen,	2.35
Ash,	1.70
	100.00

Assume that, on burning, the carbon is wholly converted into carbon dioxide and the hydrogen into water; that the nitrogen is evolved as such, and that the ash is already fully oxidized.

365. The spring water of our chalk districts is much harder than the river water of the same localities. Why is this the case?

366. How do coal and coke differ from each other in composition? What are the products obtained when (*a*) air and (*b*) steam are separately passed over red-hot coke?

367. Describe the characters and properties of six of the more important carbonates.

368. What is iron rust? How is it formed? A piece of iron is put into dilute hydrochloric acid. What happens?

369. The following are placed on a piece of sheet iron which is gradually made red-hot: sugar, starch, sal ammoniac, sulphur. What are the changes which each substance will undergo?

370. Complete the following equations:

$$MnO_2 + HCl =$$
$$NH_4Cl + CaO =$$
$$Zn + H_2SO_4 =$$
$$Zn + HCl =$$
$$NH_4NO_3 + H_2SO_4 =$$
$$NH_3 + HNO_3 =$$

371. How much nitric acid should be theoretically obtained from 1000 grams of each of the following: ammonium nitrate, potassium nitrate, sodium nitrate?

372. How would you demonstrate the presence of hydrogen in ammonia, hydrochloric acid, sulphuretted hydrogen?

373. Describe the action of heat on (*a*) iodine, (*b*) nitric acid (*c*) ammonia solution.

APPENDIX I.

RESULTS OF SOME QUANTITATIVE EXPERIMENTS.

The following are some class results of experiments described in the previous pages. They are taken at random from students' notebooks, and will afford some idea of the degree of accuracy that may be expected.

Weight of Salt in 1 litre of given solution. (30 grm. per litre.)	Mixture of 40 per cent. sand and 60 per cent. salt. (Second determination.)	Solubility of $KClO_3$ at temperature of laboratory.	Vol. of H from 1 grm. Zinc. (Uncorrected.)	Percentage of Water of Crystallization in $CuSO_4.5H_2O$.	Oxygen from 100 grm. $KClO_3$.	Weight of Oxygen in 100 grm. HgO.	Weight of Oxygen combining with 100 grm. Mg.	Weight of CO_2 from 100 grm. $CaCO_3$. (By ignition.)	Weight of CO_2 from 100 grm. $NaHCO_3$. (By action of HNO_3.)	Vol. of CO_2 from 1 grm. $NaHCO_3$. (By action of HNO_3.)	Weight of Copper turned out by 1 grm. of Zinc.	Weight of Oxygen combined with 1 grm. of Tin. (By action of HNO_3.)	Wt. of Oxygen combined with 1 grm. of Copper in CuO. (By action of HNO_3.)	Weight of Oxygen combined with 1 grm. of Copper in CuO. (By reduction.)	Wt. of Silver combined with 35.5 grm. of Chlorine in AgCl. (By reduction.)
Grms.	Per Cent. Insoluble		C.cs.	Per Cent.	Grms.	Grms.	Grms.	Grms.	Grms.	C.cs.	Grms.	Grms.	Grms.	Grms.	Grms.
28.6	41.2	5.71	353	35.2	37.0	7.2	63.3	44.2	53.1	264	.92	.27	.23	.27	109
29.6	40.6	5.67	358	36.5	38.4	7.3	63.5	43.6	51.9	256	.97	.26	.20	.27	109
28.3	44.0	5.44	360	36.1	39.2	6.9	67.0	41.9	53.0	262	1.03	.27	.23	.27	107
30.8	42.8	5.67	332	37.1	38.3	8.2	67.7	43.2	56.2	271	.91	.26	.25	.25	109
31.6	42.5	5.62	347	35.3	37.6	7.5	66.8	43.7	55.4	257	.95	.29	.24	.23	111
28.6	39.0	5.63	349	34.5	38.5	7.1	64.1	42.9	53.0	261	1.10	.28	.25	.25	109
29.0	39.2	5.73	355	34.9	41.9	7.4	65.9	45.1	52.0	268	.95	.26	.25	.27	107
28.9	41.5	5.68	361	36.9	36.9	7.6	65.2	42.7	54.1	263	.93	.27	.23	.26	108
30.0	40.7	5.61	338	34.6	38.9	7.2	65.5	43.9	52.7	274	.96	.30	.23	.23	107
29.7	41.1	5.60	353	35.7	39.4	7.6	65.0	43.7	53.3	270	.96	.26	.24	.24	109

APPENDIX II.

Apparatus required in the Chemical Laboratory:—

FOR EACH STUDENT.

Kept at the Working Bench.

1 set chemical reagents.
1 bunsen burner.
1 box test-tubes and stand.
1 test-tube brush.
2 funnels.
1 filter stand.
1 blowpipe.
1 pair of crucible tongs.
1 retort stand and 2 rings.
1 tripod stand and gauze.
1 porcelain crucible and lid.
1 pipe-clay triangle.
2 porcelain dishes.
1 wash bottle.
1 sand bath.
1 drying cone.
Stirring rods.

2. Kept at Alternate Benches.

1 pneumatic trough.
1 bee-hive shelf.
1 triangular file.

3. Kept in Store Cupboards.

1 clamp.
1 glass retort.
2 watch glasses.
1 clock glass.
2 flasks, 4 oz.
1 flask, 8 oz.
1 burette stand.
1 burette.
1 pipette.
½ litre flask.
2 beakers, 8 oz.
1 porcelain boat.
1 clay pipe.
1 iron spoon.
Rubber tubing.
Copper wire.
Rubber stoppers.
1 conical flask.
1 calcium chloride tube.
3 gas jars and cover slips.
1 deflagrating spoon.
1 apparatus for determining volume of hydrogen from metal and acid.
1 thistle funnel.
Filter papers.

FOR GENERAL USE.

Small balances. One for four students.
Boxes of weights.
More delicate balances. One or more.
Small balances for rough use, with weights.
200-c.cs. measuring cylinders. One for two students.
50-c.cs. measuring cylinders. One for two students.
Muffle furnace.
Large crucible tongs.
*1 combustion furnace.
1 small furnace with iron tube.
Large sand trays with gas burners.
Flasks, 16, 20, 40 oz.
1 or more hot air ovens.
1 or more hot water ovens.
1 foot blow-pipe and bellows.
1 hot water funnel.
*1 Carius furnace.
1 gasometer.
Large funnels.
Dropping funnels (various sizes).
Wurtz flasks.
Liebig condensers.
Agate mortars and pestles.
Wedgewood mortars and pestles.
Porcelain evaporating dishes (large sizes).
Glass gas cylinders (various sizes).
Litre flasks.
Chemical thermometers.
Desiccators.
Litmus papers.
Water baths.
Platinum crucible.
2 doz. U-tubes.
Wolff's bottles.
Drying-towers.
*Set of Hempel burettes and pipettes.
Vapour-density apparatus (V. Meyer's).
*Depression of freezing point apparatus.
*Nitrometer.
Specific gravity tubes.
Sets of special reagents.
Reagent bottles.
Wide-mouth bottles. Various sizes.
Hard and soft glass tubing.
Hard glass test-tubes, 6 in. × $\frac{3}{4}$ in.
Rubber and cork stoppers.
Cork borers.
Mercury tray.
Aspirator.
Filter pump.
Asbestos fibre.
Eudiometer.
Voltameter.
1 apparatus for determining composition of gases (SO_2 and CO_2).
1 pneumatic trough with glass side.
Crystallizing dishes.
1 barometer.
Kipp's apparatus.
1 large bell jar. Small bell jars.
Lamp glass.
1 3-tubulated glass globe.
Fire-clay crucibles.
Graphite crucibles.
Wooden models of crystals.
Round-bottom flasks.
Spring clips.
Filter papers.
Pressure tubing.
Mohr's clips.
Emery paper.
Ramsey burner. Fletcher burner.
Copper wire. Iron wire.
Magnesium wire and ribbon.
Lead turnings.
Tin foil. Zinc foil.
Granulated zinc.
Aluminum foil and wire.
Copper turnings.
Copper and platinum wire.
Copper and platinum foil.

* Apparatus marked (*) will be required only for advanced students.

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